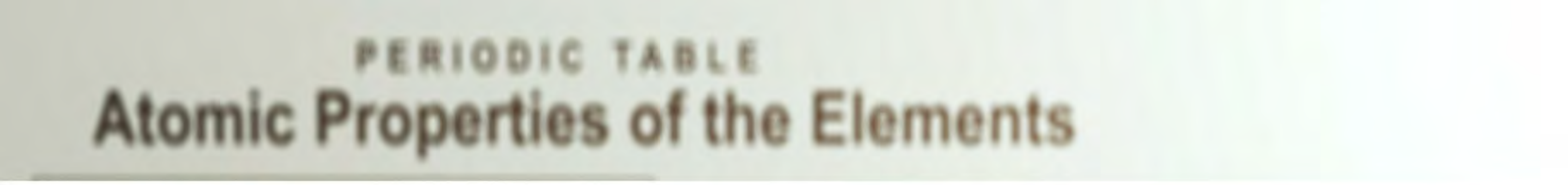

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CHEMISTRY

Fundamentals of G(ob)

Chemical Structures of Life

CHEM 100: FUNDAMENTALS OF G(OB) CHEMISTRY

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Fundamentals of GOB Chemistry (Garces)

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CHAPTER OVERVIEW

1: Chemistry Around Us

1.1: Prelude to Chemistry

If you are reading these words, you are likely starting a chemistry course. Get ready for a fantastic journey through a world of wonder, delight, and knowledge. One of the themes of this book is "chemistry is everywhere," and indeed it is; you would not be alive if it were not for chemistry, because your body is a big chemical machine.

Figure 1.1.1 © Thinkstock

6 images showing a volcanic eruption, large dinner spread, explosion, red wine, a pool of large fish, and molecules are laid out together.

If you do not believe it, do not worry. Every chapter in this book contains examples that will show you how chemistry is, in fact, everywhere. So enjoy the ride, and enjoy chemistry.

1.2: Basic Definitions

Learning Objective

- Learn the basic terms used to describe matter

The definition of chemistry—the study of the interactions of matter with other matter and with energy—uses some terms that should also be defined. We start the study of chemistry by defining basic terms.

Matter

Matter is anything that has mass and takes up space. A book is matter, a computer is matter, food is matter, and dirt in the ground is matter. Sometimes matter may be difficult to identify. For example, air is matter, but because it is so thin compared to other matter (e.g., a book, a computer, food, and dirt), we sometimes forget that air has mass and takes up space. Things that are not matter include thoughts, ideas, emotions, and hopes.

✓ Example 1.2.1

Which of the following is matter and not matter?

a. a hot dog
b. love
c. a tree

Solution

a. A hot dog has mass and takes up space, so it is matter.
b. Love is an emotion, and emotions are not matter.
c. A tree has mass and takes up space, so it is matter.

? Exercise 1.2.1

Which of the following is matter and not matter?

a. the moon
b. an idea for a new invention

Answer a

The moon is matter.

Answer b

The invention itself may be matter, but the idea for it is not.

To understand matter and how it changes, we need to be able to describe matter. There are two basic ways to describe matter: physical properties and chemical properties.

Physical properties

Physical properties are characteristics that describe matter as it exists. Some physical characteristics of matter are shape, color, size, and temperature. An important physical property is the **phase** (or **state**) of matter. The three fundamental phases of matter are solid, liquid, and gas (Figure 1.2.1).

Figure 1.2.1: The Phases of Matter. Chemistry recognizes three fundamental phases of matter: solid (left), liquid (middle), and gas (right). (CC BY-SA 3.0; Spirit469)

The solid state depicted is an ice sculpture of an angel, the liquid state is a drop of water, and the gas is clouds made of water vapor.

Chemical Properties

Chemical properties are characteristics of matter that describe how matter changes form in the presence of other matter. Does a sample of matter burn? Burning is a chemical property. Does it behave violently when put in water? This reaction is a chemical property as well (Figure 1.2.2). In the following chapters, we will see how descriptions of physical and chemical properties are important aspects of chemistry.

Figure 1.2.2: Chemical Properties. The fact that this match burns is a chemical property of the match. (Sebastian Ritter (Rise0011)).

Physical Change

A physical change occurs when a sample of matter changes one or more of its physical properties. For example, a solid may melt (Figure 1.2.3), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter.

Figure 1.2.2: Physical Changes: The solid ice melts into liquid water—a physical change. A time-lapse animation of ice cubes melting in a glass over 50 minutes. (Public Domain; Moussa).

Chemical Change

Chemical change is the process of demonstrating a chemical property, such as the burning match in Figure 1.2.2 "Chemical Properties". As the matter in the match burns, its chemical composition changes, and new forms of matter with new physical properties are created. Note that chemical changes are frequently accompanied by physical changes, as the new matter will likely have different physical properties from the original matter.

✓ Example 1.2.2

Describe each process as a physical change or a chemical change.

a. Water in the air turns into snow.
b. A person's hair is cut.
c. Bread dough becomes fresh bread in an oven.

Solution

a. Because the water is going from a gas phase to a solid phase, this is a physical change.
b. Your long hair is being shortened. This is a physical change.
c. Because of the oven's temperature, chemical changes are occurring in the bread dough to make fresh bread. These are chemical changes. (In fact, a lot of cooking involves chemical changes.)

? Exercise 1.2.2

Identify each process as a physical change or a chemical change.

a. A fire is raging in a fireplace.
b. Water is warmed to make a cup of coffee.

Answer a

chemical change

Answer b

physical change

Substance

A sample of matter that has the same physical and chemical properties throughout is called a substance. Sometimes the phrase *pure substance* is used, but the word *pure* isn't needed. The definition of the term *substance* is an example of how chemistry has a specific definition for a word that is used in everyday language with a different, vaguer definition. Here, we will use the term *substance* with its strict chemical definition.

> *Chemistry recognizes two different types of substances: elements and compounds.*

Element

An element is the simplest type of chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are 118 elements known to science, of which 80 are stable. (The other elements are radioactive, a condition we will consider in Chapter 15.) Each element has its own unique set of physical and chemical properties. Examples of elements include iron, carbon, and gold.

Compound

A compound is a combination of more than one element. The physical and chemical properties of a compound are different from the physical and chemical properties of its constituent elements; that is, it behaves as a completely different substance. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Mixtures

Physical combinations of more than one substance are called mixtures. Elements and compounds are not the only ways in which matter can be present. We frequently encounter objects that are physical combinations of more than one element or compound—mixtures. There are two types of mixtures.

Heterogeneous Mixture

A heterogeneous mixture is a mixture composed of two or more substances. It is easy to tell, sometimes by the naked eye, that more than one substance is present.

Homogeneous Mixture/ Solution

A homogeneous mixture is a combination of two or more substances that is so intimately mixed, that the mixture behaves as a single substance. Another word for a homogeneous mixture is a solution. Thus, a combination of salt and steel wool is a heterogeneous mixture because it is easy to see which particles of the matter are salt crystals and which are steel wool. On the other hand, if you take salt crystals and dissolve them in water, it is very difficult to tell that you have more than one substance present just by looking—even if you use a powerful microscope. The salt dissolved in water is a homogeneous mixture, or a solution (Figure $1.2.3$).

Figure $1.2.3$: Types of Mixtures © Thinkstock. On the left, the combination of two substances is a heterogeneous mixture because the particles of the two components look different. On the right, the salt crystals have dissolved in the water so finely that you cannot tell that salt is present. The homogeneous mixture appears like a single substance.

✓ Example $1.2.3$

Identify the following combinations as heterogeneous mixtures or homogenous mixtures.

a. soda water (carbon dioxide is dissolved in water)
b. a mixture of iron metal filings and sulfur powder (both iron and sulfur are elements)

Solution

a. Because carbon dioxide is dissolved in water, we can infer from the behavior of salt crystals dissolved in water that carbon dioxide dissolved in water is (also) a homogeneous mixture.
b. Assuming that the iron and sulfur are simply mixed together, it should be easy to see what is iron and what is sulfur, so this is a heterogeneous mixture.

? Exercise $1.2.3$

a. the human body
b. an amalgam, a combination of some other metals dissolved in a small amount of mercury

Answer a

heterogeneous mixture

Answer b

homogeneous mixture

There are other descriptors that we can use to describe matter, especially elements. We can usually divide elements into metals and nonmetals, and each set shares certain (but not always all) properties.

Metal

A metal is an element that conducts electricity and heat well and is shiny, silvery, solid, ductile, and malleable. At room temperature, metals are solid (although mercury is a well-known exception). A metal is ductile because it can be drawn into thin wires (a property called *ductility*); and malleable because it can be pounded into thin sheets (a property called *malleability*).

Nonmetal

A non-metal is an element that is brittle when solid, and does not conduct electricity or heat very well. Non-metals cannot be made into thin sheets or wires (Figure 1.2.4). Nonmetals also exist in a variety of phases and colors at room temperature.

Semi-metals

Some elements have properties of both metals and nonmetals and are called semi-metals (or metalloids). We will see later how these descriptions can be assigned rather easily to various elements.

Figure 1.2.4: Semimetals © Thinkstock. On the left is some elemental mercury, the only metal that exists as a liquid at room temperature. It has all the other expected properties of a metal. On the right, elemental sulfur is a yellow nonmetal that usually is found as a powder.

Describing Matter Flowchart

"Describing Matter" is a flowchart of the relationships among the different ways of describing matter.

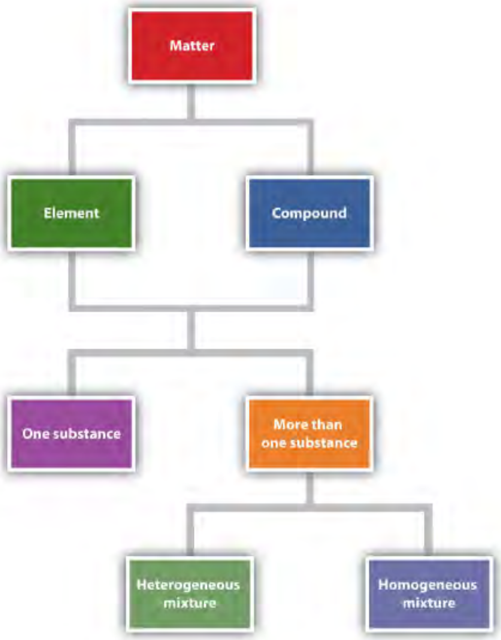


Figure 1.2.5: Describing Matter. This flowchart shows how matter can be described.

Matter forks into element & compound which fork into 1 substance & more than 1 substance. More than 1 substance is homogenous or heterogenous.

✓ Example 1.2.1: Chemistry is Everywhere: In the Morning

Most people have a morning ritual, a process that they go through every morning to get ready for the day. Chemistry appears in many of these activities.

- If you take a shower or bath in the morning, you probably use soap, shampoo, or both. These items contain chemicals that interact with the oil and dirt on your body and hair to remove them and wash them away. Many of these products also contain chemicals that make you smell good; they are called *fragrances*.
- When you brush your teeth in the morning, you usually use toothpaste, a form of soap, to clean your teeth. Toothpastes typically contain tiny, hard particles called *abrasives* that physically scrub your teeth. Many toothpastes also contain fluoride, a substance that chemically interacts with the surface of the teeth to help prevent cavities.
- Perhaps you take vitamins, supplements, or medicines every morning. Vitamins and other supplements contain chemicals your body needs in small amounts to function properly. Medicines are chemicals that help combat diseases and promote health.
- Perhaps you make some fried eggs for breakfast. Frying eggs involves heating them enough so that a chemical reaction occurs to cook the eggs.
- After you eat, the food in your stomach is chemically reacted so that the body (mostly the intestines) can absorb food, water, and other nutrients.
- If you drive or take the bus to school or work, you are using a vehicle that probably burns gasoline, a material that burns fairly easily and provides energy to power the vehicle. Recall that burning is a chemical change.

These are just a few examples of how chemistry impacts your everyday life. And we haven't even made it to lunch yet!

Figure 1.2.6: Chemistry in Real Life © Thinkstock. Examples of chemistry can be found everywhere—in personal hygiene products, food, and motor vehicles.

Personal hygiene products (left), food (middle), and motor vehicles (right) images next to one another.

Key Takeaways

- Chemistry is the study of matter and its interactions with other matter and energy.
- Matter is anything that has mass and takes up space.
- Matter can be described in terms of physical properties and chemical properties.
- Physical properties and chemical properties of matter can change.
- Matter is composed of elements and compounds.
- Combinations of different substances are called mixtures.
- Elements can be described as metals, nonmetals, and semi-metals.

1.3: What is Chemistry?

Learning Objectives

1. Define chemistry in relation to other sciences.
2. Identify the general steps in the scientific method.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Being able to describe the ingredients in a cake and how they change when the cake is baked is called chemistry. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter.

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure $1.3.1$). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

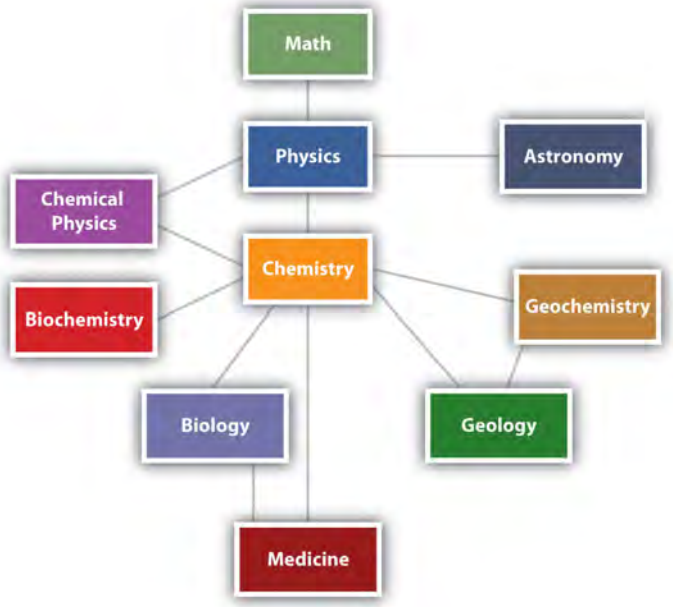


Figure $1.3.1$: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

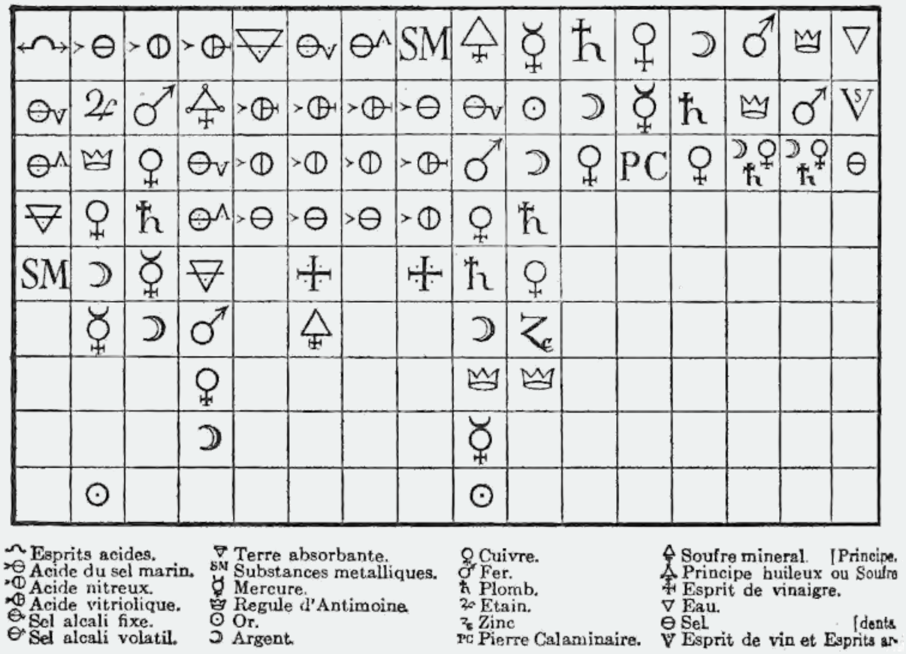
Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure $1.3.1$ shows how many of the individual fields of science are related.

> *There are many other fields of science in addition to the ones listed here.*

Alchemy

As our understanding of the universe has changed over time, so has the practice of science. Chemistry in its modern form, based on principles that we consider valid today, was developed in the 1600s and 1700s. Before that, the study of matter was known as *alchemy* and was practiced mainly in China, Arabia, Egypt, and Europe.

Alchemy was a somewhat mystical and secretive approach to learning how to manipulate matter. Practitioners, called alchemists, thought that all matter was composed of different proportions of the four basic elements—fire, water, earth, and air—and believed that if you changed the relative proportions of these elements in a substance, you could change the substance. The long-standing attempts to "transmute" common metals into gold represented one goal of alchemy. Alchemy's other major goal was to synthesize the philosopher's stone, a material that could impart long life—even immortality. Alchemists used symbols to represent substances, some of which are shown in the accompanying figure. This was not done to better communicate ideas, as chemists do today, but to maintain the secrecy of alchemical knowledge, keeping others from sharing in it.



The first affinity table. Table of different relations observed in chemistry between different substances; Memoirs of the Royal Academy of Sciences, p. 202-212. Alchemists used symbols like these to represent substances.

In spite of this secrecy, in its time alchemy was respected as a serious, scholarly endeavor. Isaac Newton, the great mathematician and physicist, was also an alchemist.

✓ Exercise 1.3.1

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Answer a

Sculpture is not considered a science because it is not a study of some aspect of the natural universe.

Answer b

Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

? Exercise 1.3.2

Which fields of study are branches of science?

a. physiology (the study of the function of an animal's or a plant's body)
b. geophysics
c. agriculture
d. politics

Answer

a, b and c only

How do scientists work? Generally, they follow a process called the scientific method. The scientific method is an organized procedure for learning answers to questions and making explanations for observations. To find the answer to a question (for example, “Why do birds fly toward Earth’s equator during the cold months?”), a scientist goes through the following steps, which are also illustrated in Figure $1.3.2$:

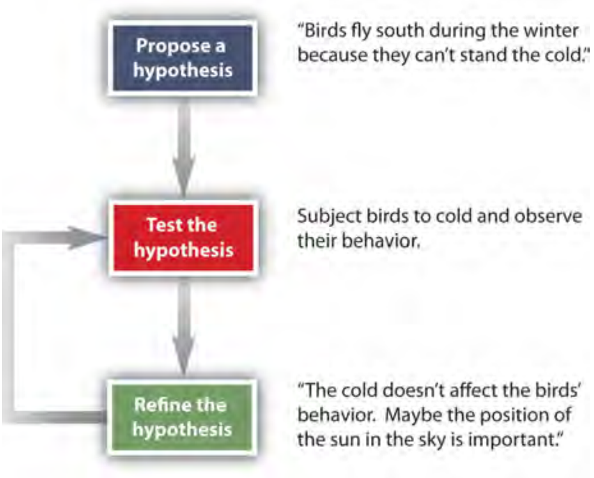


Figure $1.3.2$: The General Steps of the Scientific Method. After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.

The steps may not be as clear-cut in real life as described here, but most scientific work follows this general outline.

1. **Propose a hypothesis.** A scientist generates a testable idea, or hypothesis, to try to answer a question or explain an observation about how the natural universe works. Some people use the word *theory* in place of hypothesis, but the word hypothesis is the proper word in science. For scientific applications, the word theory is a general statement that describes a large set of observations and data. A theory represents the highest level of scientific understanding.
2. **Test the hypothesis.** A scientist evaluates the hypothesis by devising and carrying out experiments to test it. If the hypothesis passes the test, it may be a proper answer to the question. If the hypothesis does not pass the test, it may not be a good answer.
3. **Refine the hypothesis if necessary.** Depending on the results of experiments, a scientist may want to modify the hypothesis and then test it again. Sometimes the results show the original hypothesis to be completely wrong, in which case a scientist will have to devise a new hypothesis.

Not all scientific investigations are simple enough to be separated into these three discrete steps. But these steps represent the general method by which scientists learn about our natural universe.

– rcise

Define science and chemistry.

Answer

Science is a process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Chemistry is the study of matter.

– rcise

Name the steps of the scientific method.

Answer

After identifying the problem or making an observation, propose a hypothesis, test the hypothesis, and refine the hypothesis, if necessary

Key Takeaways

- Chemistry is the study of matter and how it behaves.
- The scientific method is the general process by which we learn about the natural universe.

1. Based on what you know, which fields are branches of science?
    a. meteorology (the study of weather)
    b. astrophysics (the physics of planets and stars)
    c. economics (the study of money and monetary systems)
    d. astrology (the prediction of human events based on planetary and star positions)
    e. political science (the study of politics)
2. Based on what you know, which fields are a branches of science?
    a. history (the study of past events)
    b. ornithology (the study of birds)
    c. paleontology (the study of fossils)
    d. zoology (the study of animals)
    e. phrenology (using the shape of the head to determine personal characteristics)
3. Which of the following are examples of matter?
    a. a baby
    b. an idea
    c. the Empire State Building
    d. an emotion
    e. the air
    f. Alpha Centauri, the closest known star (excluding the sun) to our solar system
4. Which of the following are examples of matter?
    a. your textbook
    b. brain cells
    c. love
    d. a can of soda
    e. breakfast cereal
5. Suggest a name for the science that studies the physics of rocks and the earth.
6. Suggest a name for the study of the physics of living organisms.
7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.
8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.
9. When someone says, “I have a theory that excess salt causes high blood pressure,” does that person really have a theory? If it is not a theory, what is it?
10. When a person says, “My hypothesis is that excess calcium in the diet causes kidney stones,” what does the person need to do to determine if the hypothesis is correct?
11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?
12. Most students take multiple English classes in school. Does the study of English use the scientific method?

Answers

1.
    a. science
    b. science

c. not science
d. not science
e. not science

2.

a. not science
b. science
c. science
d. science
e. not science

3.

a. matter
b. not matter
c. matter
d. not matter
e. matter
f. matter

4.

a. matter
b. matter
c. not matter
d. matter
e. matter

5. geophysics

6. biophysics

7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.

8. Engineering

9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

1.4: Chemicals Compose Ordinary Things

Chemistry is the branch of science dealing with the structure, composition, properties, and the reactive characteristics of matter. Matter is anything that has mass and occupies space. Thus, chemistry is the study of literally everything around us—the liquids that we drink, the gases we breathe, the composition of everything from the plastic case on your phone to the earth beneath your feet. Moreover, chemistry is the study of the transformation of matter. Crude oil is transformed into more useful petroleum products, such as gasoline and kerosene, by the process of refining. Some of these products are further transformed into plastics. Crude metal ores are transformed into metals, that can then be fashioned into everything from foil to automobiles. Potential drugs are identified from natural sources, isolated and then prepared in the laboratory. Their structures are systematically modified to produce the pharmaceuticals that have led to vast advances in modern medicine. Chemistry is at the center of all of these processes; chemists are the people that study the nature of matter and learn to design, predict, and control these chemical transformations. Within the branches of chemistry you will find several apparent subdivisions. Inorganic chemistry, historically, focused on minerals and metals found in the earth, while organic chemistry dealt with carbon-containing compounds that were first identified in living things. Biochemistry is an outgrowth of the application of organic chemistry to biology and relates to the chemical basis for living things. In the later chapters of this text we will explore organic and biochemistry in a bit more detail and you will notice examples of organic compounds scattered throughout the text. Today, the lines between the various fields have blurred significantly and a contemporary chemist is expected to have a broad background in all of these areas.

In this chapter, we will discuss some of the properties of matter and how chemists measure those properties. We will introduce some of the vocabulary that is used throughout chemistry and the other physical sciences.

Let's begin with matter. **Matter** is defined as any substance that has mass. It is important to distinguish here between weight and mass. Weight is the result of the pull of gravity on an object. On the Moon, an object will weigh less than the same object on Earth because the pull of gravity is less on the Moon. The mass of an object, however, is an inherent property of that object and does not change, regardless of location, gravitational pull, or anything else. It is a property that is solely dependent on the quantity of matter within the object.

Contemporary theories suggests that matter is composed of **atoms**. Atoms themselves are constructed from neutrons, protons and electrons, along with an ever-increasing array of other subatomic particles. We will focus on the neutron, a particle having no charge; the proton, which carries a positive charge; and the electron, which has a negative charge. Atoms are incredibly small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that's 28 sextillion). Because atoms and subatomic particles are so small, their mass is not readily measured using pounds, ounces, grams or any other scale that we would use on larger objects. Instead, the mass of atoms and subatomic particles is measured using atomic mass units (abbreviated amu). The atomic mass unit is based on a scale that relates the mass of different types of atoms to each other (using the most common form of the element carbon as a standard). The amu scale gives us a convenient means to describe the masses of individual atoms and to do quantitative measurements concerning atoms and their reactions. Within an atom, the neutron and proton both have a mass of one amu; the electron has a much smaller mass (about 0.0005 amu).

Figure 1.4.1: Atoms are incredible small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that's 28 sextillion).

Atomic theory places the neutron and the proton in the center of the atom in the nucleus. In an atom, the nucleus is very small, very dense, carries a positive charge (from the protons) and contains virtually all of the mass of the atom. Electrons are placed in a diffuse cloud surrounding the nucleus. The electron cloud carries a net negative charge (from the charge on the electrons) and in a neutral atom there are always as many electrons in this cloud as there are protons in the nucleus (the positive charges in the nucleus are balanced by the negative charges of the electrons, making the atom neutral).

An atom is characterized by the number of neutrons, protons and electrons that it possesses. Today, we recognize at least 116 different types of atoms, each type having a different number of protons in its nucleus. These different types of atoms are called elements. The neutral element hydrogen (the lightest element) will always have one proton in its nucleus and one electron in the cloud surrounding the nucleus. The element helium will always have two protons in its nucleus. It is the number of protons in the nucleus of an atom that defines the identity of an element. Elements can, however, have differing numbers of neutrons in their nucleus. For example, stable helium nuclei exist that contain one, or two neutrons (but they all have two protons). These different types of helium atoms have different masses (3 or 4 amu) and they are called isotopes. For any given isotope, the sum of the numbers of protons and neutrons in the nucleus is called the mass number. All elements exist as a collection of isotopes, and the mass of an element that we use in chemistry, the atomic mass, is the average of the masses of these isotopes. For helium, there is approximately one isotope of Helium-3 for every one million isotopes of Helium-4, hence the average atomic mass is very close to 4 (4.002602).

As different elements were discovered and named, abbreviations of their names were developed to allow for a convenient chemical shorthand. The abbreviation for an element is called its chemical symbol. A chemical symbol consists of one or two letters, and the relationship between the symbol and the name of the element is generally apparent. Thus helium has the chemical symbol He, nitrogen is N, and lithium is Li. Sometimes the symbol is less apparent but is decipherable; magnesium is Mg, strontium is Sr, and manganese is Mn. Symbols for elements that have been known since ancient times, however, are often based on Latin or Greek names and appear somewhat obscure from their modern English names. For example, copper is Cu (from cuprum), silver is Ag (from argentum), gold is Au (from aurum), and iron is Fe (from ferrum). Throughout your study of chemistry, you will routinely use chemical symbols and it is important that you begin the process of learning the names and chemical symbols for the common elements. By the time you complete General Chemistry, you will find that you are adept at naming and identifying virtually all of the 116 known elements. Table $1.4.1$ contains a starter list of common elements that you should begin learning now!

Table $1.4.1$: Names and Chemical Symbols for Common Elements

Element	Chemical Symbol	Element	Chemical Symbol
Hydrogen	H	Phosphorus	P
Helium	He	Sulfur	S
Lithium	Li	Chlorine	Cl
Beryllium	Be	Argon	Ar
Boron	B	Potassium	K
Carbon	C	Calcium	Ca
Nitrogen	N	Iron	Fe
Oxygen	O	Copper	Cu
Fluorine	F	Zinc	Zn
Neon	Ne	Bromine	Br
Sodium	Na	Silver	Ag
Magnesium	Mg	Iodine	I
Aluminum	Al	Gold	Au
Silicon	Si	Lead	Pb

The chemical symbol for an element is often combined with information regarding the number of protons and neutrons in a particular isotope of that atom to give the atomic symbol. To write an atomic symbol, begin with the chemical symbol, then write the atomic number for the element (the number of protons in the nucleus) as a subscript, preceding the chemical symbol. Directly above this, as a superscript, write the mass number for the isotope, that is, the total number of protons and neutrons in the nucleus. Thus, for helium, the atomic number is 2 and there are two neutrons in the nucleus for the most common isotope, making the atomic symbol $^{4}_{2}\text{He}$. In the definition of the atomic mass unit, the “most common isotope of carbon”, $^{12}_{6}\text{C}$, is defined as having a mass of exactly 12 amu and the atomic masses of the remaining elements are based on their masses relative to this isotope. Chlorine (chemical symbol Cl) consists of two major isotopes, one with 18 neutrons (the most common, comprising 75.77% of natural chlorine atoms) and one with 20 neutrons (the remaining 24.23%). The atomic number of chlorine is 17 (it has 17 protons in its nucleus), therefore the chemical symbols for the two isotopes are $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$.

When data is available regarding the natural abundance of various isotopes of an element, it is simple to calculate the average atomic mass. In the example above, $^{35}_{17}\text{Cl}$ was the most common isotope with an abundance of 75.77% and $^{37}_{17}\text{Cl}$ had an abundance of the remaining 24.23%. To calculate the average mass, first convert the percentages into fractions; that is, simply divide them by 100. Now, chlorine-35 represents a fraction of natural chlorine of 0.7577 and has a mass of 35 (the mass number). Multiplying these, we get (0.7577 × 35) = 26.51. To this, we need to add the fraction representing chlorine-37, or (0.2423 × 37) = 8.965; adding, (26.51 + 8.965) = 35.48, which is the weighted average atomic mass for chlorine. Whenever we do mass calculations involving elements or compounds (combinations of elements), we always need to use average atomic masses.

Contributions & Attributions

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1.5: Hypothesis, Theories, and Laws

Learning Objectives

- Describe the difference between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.

Although many have taken science classes throughout the course of their studies, people often have incorrect or misleading ideas about some of the most important and basic principles in science. Most students have heard of hypotheses, theories, and laws, but what do these terms really mean? Prior to reading this section, consider what you have learned about these terms before. What do these terms mean to you? What do you read that contradicts or supports what you thought?

What is a Fact?

A fact is a basic statement established by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

One of the most common terms used in science classes is a "hypothesis". The word can have many different definitions, depending on the context in which it is being used:

- An educated guess: a scientific hypothesis provides a suggested solution based on evidence.
- Prediction: if you have ever carried out a science experiment, you probably made this type of hypothesis when you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for it to be scientific, however, a scientist must be able to test the explanation to see if it works and if it is able to correctly predict what will happen in a situation. For example, "if my hypothesis is correct, we should see ___ result when we perform ___ test."

> *A hypothesis is very tentative; it can be easily changed.*

What is a Theory?

The *United States National Academy of Sciences* describes what a theory is as follows:

> "Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."
>
> "A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact.

Note some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomena. They aren't predictions (although we may use theories to make predictions). They are explanations as to why we observe something.
- Theories aren't likely to change. They have a large amount of support and are able to satisfactorily explain numerous observations. Theories can, indeed, be facts. Theories can change, but it is a long and difficult process. In order for a theory to change, there must be many observations or pieces of evidence that the theory cannot explain.
- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight; it is not just one person's idea about something

> *Theories aren't likely to change.*

What is a Law?

Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are more overarching explanations of how nature works and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas is increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

What is a Belief?

A **belief** is a statement that is not scientifically provable. Beliefs may or may not be incorrect; they just are outside the realm of science to explore.

Laws vs. Theories

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, **theories are explanations** and **laws are patterns** we see in large amounts of data, frequently written as an equation. A theory will always remain a theory; a law will always remain a law.

Video $1.5.1$: What’s the difference between a scientific law and theory?

Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.

Contributions & Attributions

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- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

1.6: The Scientific Method - How Chemists Think

Learning Objectives

- Identify the components of the scientific method.

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments; which leads to additional observations, hypotheses, and experiments in repeated cycles (Figure $1.6.1$).

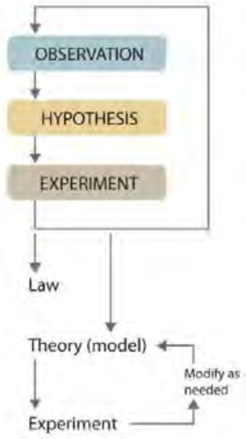

Figure $1.6.1$: The Steps in the Scientific Method.

Step 1: Make observations

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: "the outside air temperature is cooler during the winter season," "table salt is a crystalline solid," "sulfur crystals are yellow," and "dissolving a penny in dilute nitric acid forms a blue solution and a brown gas." **Quantitative observations** are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: "the melting point of crystalline sulfur is 115.21° Celsius," and "35.9 grams of table salt—the chemical name of which is sodium chloride—dissolve in 100 grams of water at 20° Celsius." For the question of the dinosaurs' extinction, the initial observation was quantitative: iridium concentrations in sediments dating to 66 million years ago were 20–160 times higher than normal.

Step 2: Formulate a hypothesis

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a hypothesis, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience alternating periods of light and darkness corresponding to observed movements of the sun, moon, clouds, and shadows is consistent with either one of two hypotheses:

a. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
b. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either support or refute it.

Step 3: Design and perform experiments

After a hypothesis has been formed, scientists conduct experiments to test its validity. Experiments are systematic observations or measurements, preferably made under controlled conditions—that is—under conditions in which a single variable changes.

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether or not the original hypothesis is valid. If the hypothesis is valid, the scientist can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified and requires further experimentation.

Step 5: Development into a law and/or theory

More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a law, a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply states what happens; it does not address the question of why.

One example of a law, the law of definite proportions, which was discovered by the French scientist Joseph Proust (1754–1826), states that a chemical substance always contains the same proportions of elements by mass. Thus, sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine, in this case 39.34% sodium and 60.66% chlorine by mass, and sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass.

Whereas a law states only what happens, a theory attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time unless a major experimental error is discovered. In contrast, a theory, by definition, is incomplete and imperfect, evolving with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure 1.6.1 at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

✓ Example 1.6.1

Classify each statement as a law, a theory, an experiment, a hypothesis, an observation.

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
d. When 10 g of ice were added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
e. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is a possible explanation for the origin of birds, so it is a hypothesis.
c. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
d. The temperature is measured before and after a change is made in a system, so these are observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

? Exercise 1.6.1

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter to ever play professional basketball.
e. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

observation

Summary

The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. The key steps in the scientific method include the following:

- Step 1: Make observations.
- Step 2: Formulate a hypothesis.
- Step 3: Test the hypothesis through experimentation.
- Step 4: Accept or modify the hypothesis.
- Step 5: Develop into a law and/or a theory.

Contributions & Attributions

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- Wikipedia

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1.7: Learning Chemistry

Learning Objective

- Learn what science is and how it works.

Chemistry is a branch of science. Although science itself is difficult to define exactly, the following definition can serve as a starting point. Science is the process of knowing about the natural universe through observation and experiment. Science is not the only process of knowing (e.g., the ancient Greeks simply sat and *thought*), but it has evolved over more than 350 years into the best process that humanity has devised, to date, to learn about the universe around us.

The process of science is usually stated as the *scientific method*, which is rather naively described as follows:

1. state a hypothesis,
2. test the hypothesis, and
3. refine the hypothesis

In actuality, the process is not that simple. (For example, a scientist does not go into their lab every day and exclaim, "I am going to state a hypothesis today and spend the day testing it!") The process is not that simple because science and scientists have a body of knowledge that has already been identified as coming from the highest level of understanding, and most scientists build from that body of knowledge.

An educated guess about how the natural universe works is called a **hypothesis**. A scientist who is familiar with how part of the natural universe works—say, a chemist—is interested in furthering that knowledge. That person makes a reasonable guess—a hypothesis—that is designed to see if the universe works in a new way as well. Here's an example of a hypothesis: "if I mix one part of hydrogen with one part of oxygen, I can make a substance that contains both elements."

> *For a hypothesis to be termed a scientific hypothesis, it has to be something that can be supported or refuted through carefully crafted experimentation or observation.*

Most good hypotheses are grounded in previously understood knowledge and represent a testable extension of that knowledge. The scientist then devises ways to test if that hypothesis is correct or not. That is, the scientist plans experiments. **Experiments are tests of the natural universe to see if a guess (hypothesis) is correct**. An experiment to test our previous hypothesis would be to actually mix hydrogen and oxygen and see what happens. Most experiments include observations of small, well-defined parts of the natural universe designed to see results of the experiments.

A Scientific Hypothesis

A hypothesis is often written in the form of an if/then statement that gives a possibility (if) and explains what may happen because of the possibility (then). For example, if eating elemental sulfur repels ticks, then someone that is eating sulfur every day will not get ticks.

Why do we have to do experiments? Why do we have to test? Because the natural universe is not always so obvious, experiments are necessary. For example, it is fairly obvious that if you drop an object from a height, it will fall. Several hundred years ago (coincidentally, near the inception of modern science), the concept of gravity explained that test. However, is it obvious that the entire natural universe is composed of only about 115 fundamental chemical building blocks called elements? This wouldn't seem true if you looked at the world around you and saw all the different forms matter can take. In fact, the concept of *the element* is only about 200 years old, and the last naturally occurring element was identified about 80 years ago. It took decades of tests and millions of experiments to establish what the elements actually are. These are just two examples; a myriad of such examples exists in chemistry and science in general.

When enough evidence has been collected to establish a general principle of how the natural universe works, the evidence is summarized in a theory. A theory is a general statement that explains a large number of observations. "All matter is composed of atoms" is a general statement, a theory, that explains many observations in chemistry. A theory is a very powerful statement in science. There are many statements referred to as "the theory of _______" or the "______ theory" in science (where the blanks represent a word or concept). When written in this way, theories indicate that science has an overwhelming amount of evidence of its correctness. We will see several theories in the course of this text.

A specific statement that is thought to never be violated by the entire natural universe is called a law. A scientific law is the highest understanding of the natural universe that science has and is thought to be inviolate. The fact that all matter attracts all other matter—the law of gravitation—is one such law. Note that the terms *theory* and *law* used in science have slightly different meanings from those in common usage; where theory is often used to mean hypothesis ("I have a theory…"), and a law is an arbitrary limitation that can be broken but with potential consequences (such as speed limits). Here again, science uses these terms differently, and it is important to apply their proper definitions when you use these words in science. (Figure 1.7.1)

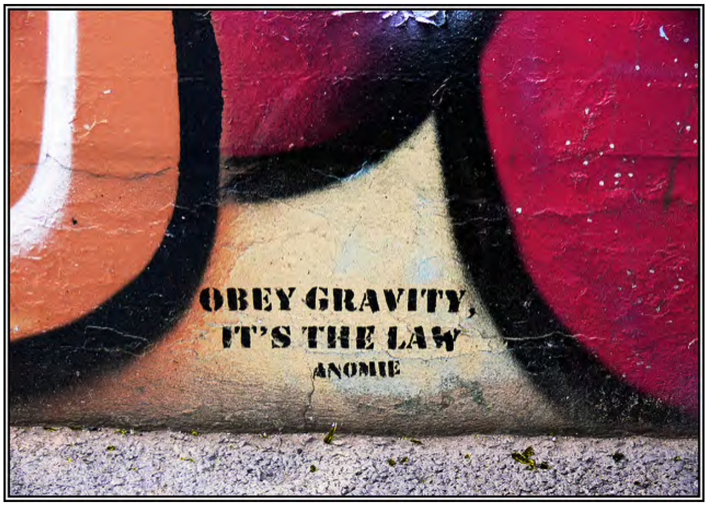

Figure 1.7.1: Defining a law. Does this graffiti mean "law" the way science defines "law"? (CC BY-SA-NC-ND; Koppenbadger).

There is an additional phrase in our definition of science: "the natural universe." Science is concerned *only* with the natural universe. What is the natural universe? It's anything that occurs around us, well, naturally. Stars, planets, the appearance of life on earth; as well as how animals, plants, and other matter function are all part of the natural universe. Science is concerned with that—and *only* that.

Of course, there are other things that concern us. For example, is the English language part of science? Most of us can easily answer no; English is not science. English is certainly worth knowing (at least for people in predominantly English-speaking countries), but why isn't it science? English, or any human language, is not science because ultimately it is *contrived*; it is made up. Think of it: the word spelled b-l-u-e represents a certain color, and we all agree what color that is. But what if we used the word h-a-r-d to describe that color? (Figure 1.7.2) That would be fine—as long as everyone agreed. Anyone who has learned a second language must initially wonder why a certain word is used to describe a certain concept; ultimately, the speakers of that language agreed that a particular word would represent a particular concept. It was contrived.

That doesn't mean language isn't worth knowing. It is very important in society. But it's not *science*. Science deals only with what occurs naturally.

Figure 1.7.2: English Is Not Science. How would you describe this color? Blue or hard? Either way, you're not doing science.

✓ Example 1.7.1: Identifying Science

Which of the following fields would be considered science?

a. geology, the study of the earth
b. ethics, the study of morality
c. political science, the study of governance
d. biology, the study of living organisms

Solution

a. Because the earth is a natural object, the study of it is indeed considered part of science.
b. Ethics is a branch of philosophy that deals with right and wrong. Although these are useful concepts, they are not science.
c. There are many forms of government, but all are created by humans. Despite the fact that the word *science* appears in its name, political science is not true science.
d. Living organisms are part of the natural universe, so the study of them is part of science.

? Exercise 1.7.1

Which is part of science, and which is not?

a. dynamics, the study of systems that change over time
b. aesthetics, the concept of beauty

Answer A

science

Answer B

not science

The field of science has gotten so big that it is common to separate it into more specific fields. First, there is mathematics, the language of science. All scientific fields use mathematics to express themselves—some more than others. Physics and astronomy are scientific fields concerned with the fundamental interactions between matter and energy. Chemistry, as defined previously, is the study of the interactions of matter with other matter and with energy. Biology is the study of living organisms, while geology is the study of the earth. Other sciences can be named as well. Understand that these fields are not always completely separate; the boundaries between scientific fields are not always readily apparent. A scientist may be labeled a biochemist if he or she studies the chemistry of biological organisms.

Finally, understand that science can be either qualitative or quantitative. Qualitative implies a description of the quality of an object. For example, physical properties are generally qualitative descriptions: sulfur is yellow, your math book is heavy, or that statue is pretty. A quantitative description represents the specific amount of something; it means knowing how much of something is present, usually by counting or measuring it. Some quantitative descriptions include: 25 students in a class, 650 pages in a book, or a velocity of 66 miles per hour. Quantitative expressions are very important in science; they are also very important in chemistry.

✓ Example 1.7.2: qualitative vs. quantitative Descriptions

Identify each statement as either a qualitative description or a quantitative description.

a. Gold metal is yellow.
b. A ream of paper has 500 sheets in it.
c. The weather outside is snowy.
d. The temperature outside is 24 degrees Fahrenheit.

Solution

a. Because we are describing a physical property of gold, this statement is qualitative.
b. This statement mentions a specific amount, so it is quantitative.
c. The word *snowy* is a description of how the day is; therefore, it is a qualitative statement.
d. In this case, the weather is described with a specific quantity—the temperature. Therefore, it is quantitative.

? Exercise 1.7.2

Are these qualitative or quantitative statements?

a. Roses are red, and violets are blue.
b. Four score and seven years ago….

Answer A

qualitative

Answer B

quantitative

Food and Drink Application: Carbonated Beverages

Some of the simple chemical principles discussed in this chapter can be illustrated with carbonated beverages: sodas, beer, and sparkling wines. Each product is produced in a different way, but they all have one thing in common: they are solutions of carbon dioxide dissolved in water.

Carbon dioxide is a compound composed of carbon and oxygen. Under normal conditions, it is a gas. If you cool it down enough, it becomes a solid known as dry ice. Carbon dioxide is an important compound in the cycle of life on earth.

Even though it is a gas, carbon dioxide can dissolve in water, just like sugar or salt can dissolve in water. When that occurs, we have a homogeneous mixture, or a solution, of carbon dioxide in water. However, very little carbon dioxide can dissolve in water. If the atmosphere were pure carbon dioxide, the solution would be only about 0.07% carbon dioxide. In reality, the air is only about 0.03% carbon dioxide, so the amount of carbon dioxide in water is reduced proportionally.

However, when soda and beer are made, manufacturers do two important things: they use pure carbon dioxide gas, and they use it at very high pressures. With higher pressures, more carbon dioxide can dissolve in the water. When the soda or beer container is sealed, the high pressure of carbon dioxide gas remains inside the package. (Of course, there are more ingredients in soda and beer besides carbon dioxide and water.)

When you open a container of soda or beer, you hear a distinctive *hiss* as the excess carbon dioxide gas escapes. But something else happens as well. The carbon dioxide in the solution comes out of solution as a bunch of tiny bubbles. These bubbles impart a pleasing sensation in the mouth, so much so that the soda industry sold over *225 billion* servings of soda in the United States alone in 2009.

Some sparkling wines are made in the same way—by forcing carbon dioxide into regular wine. Some sparkling wines (including champagne) are made by sealing a bottle of wine with some yeast in it. The yeast *ferments*, a process by which the yeast converts sugars into energy and excess carbon dioxide. The carbon dioxide produced by the yeast dissolves in the wine. Then, when the champagne bottle is opened, the increased pressure of carbon dioxide is released, and the drink bubbles just like an expensive glass of soda.

Figure 1.7.3: Carbonated Beverages © Thinkstock. Soda, beer, and sparkling wine take advantage of the properties of a solution of carbon dioxide in water.

Soda (left), beer (middle) and sparkling wine (right) images next to one another.

Key Takeaways

- Science is a process of knowing about the natural universe through observation and experiment.
- Scientists go through a rigorous process to determine new knowledge about the universe; this process is generally referred to as the scientific method.
- Science is broken down into various fields, of which chemistry is one.
- Science, including chemistry, is both qualitative and quantitative.

1.8: A Beginning Chemist - How to Succeed

Examples of the practical applications of chemistry are everywhere (Figure 1.8.1). Engineers need to understand the chemical properties of the substances needed to design biologically compatible implants for joint replacements; or to design roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Archeology and paleontology rely on chemical techniques to date bones and artifacts and identify their origins. Although law is not normally considered a field related to chemistry, forensic scientists use chemical methods to analyze blood, fibers, and other evidence as they investigate crimes. In particular, DNA matching—comparing biological samples of genetic material to see whether they could have come from the same person—has been used to solve many high-profile criminal cases as well as clear innocent people who have been wrongly accused or convicted. Forensics is a rapidly growing area of applied chemistry. In addition, the proliferation of chemical and biochemical innovations in industry is producing rapid growth in the area of patent law. Ultimately, the dispersal of information in all the fields in which chemistry plays a part requires experts who are able to explain complex chemical issues to the public through television, print journalism, the Internet, and popular books.

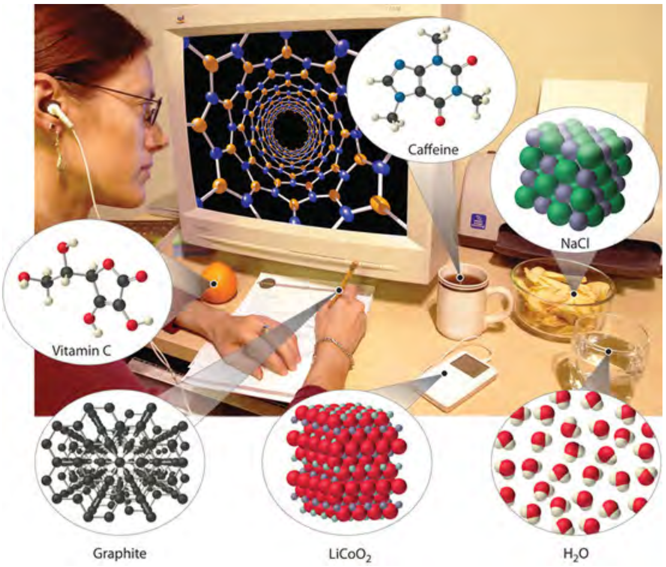

Figure 1.8.1: Chemistry in Everyday Life. Although most people do not recognize it, chemistry and chemical compounds are crucial ingredients in almost everything we eat, wear, and use.

Chemical compounds in everyday life: Vitamin C, graphite, lithium cobalt oxide, caffeine, sodium chloride, water

Hopefully at this point you are fully convinced of how important and useful the study of chemistry can be. You may, however, still be wondering exactly *what* it is that a chemist does. Chemistry is the study of matter and the changes that matter undergoes. In general, chemists are interested in both characteristics that you can test and observe, like a chemical's smell or color, and characteristics that are far too small to see, like what the oxygen you breathe in or the carbon dioxide you breath out looks like under a microscope 1,000 times more powerful than any existing in the world today.

Wait a minute… how can a chemist *know* what oxygen and carbon dioxide look like under a microscope that doesn't even exist? What happened to the scientific method? What happened to relying on *observations* and *careful measurements*? In fact, because chemists *can't* see the underlying structure of different materials, they have to rely on the scientific method even more! Chemists are a lot like detectives. Suppose a detective is trying to solve a murder case—what do they do? Obviously, the detective starts by visiting the site of the crime and looking for evidence. If the murderer has left enough clues behind, the detective can piece together a theory explaining what happened.

Even though the detective wasn't at the crime scene when the crime was committed and didn't actually see the murderer kill the victim, with the right evidence, the detective can be pretty sure of how the crime took place. It is the same with chemistry. When chemists go into the laboratory, they collect evidence by making measurements. Once chemists have collected enough clues from the properties that they can observe, they use that evidence to piece together a theory explaining the properties that they cannot observe—the properties that are too small to see.

What kinds of properties do chemists actually measure in the laboratory? Well, you can probably guess a few. Imagine that you go to dinner at a friend's house and are served something that you don't recognize, what types of observations might you make to determine exactly *what* you've been given? You might smell the food. You might note the color of the food. You might try to decide whether the food is a liquid or a solid because if it's a liquid, it's probably soup or a drink. The temperature of the food could be useful if you wanted to know whether or not you had been served ice cream! You could also pick up a small amount of food with your fork and try to figure out how much it weighs—a light dessert might be something like an angel cake, while a heavy dessert is probably a pound cake. The quantity of food you have been given might be a clue too. Finally, you might want to know something about the food's texture—is it hard and granular like sugar cubes, or soft and easy to spread, like butter?

Believe it or not, the observations you are likely to make when trying to identify an unknown food are very similar to the observations that a chemist makes when trying to learn about a new material. Chemists rely on smell, color, state (whether it is a solid or liquid or gas), temperature, volume, mass (which is related to weight—as will be discussed in a later section), and texture. There is, however, one property possibly used to learn about a food, but that should *definitely* not be used to learn about a chemical—taste!

In the sections on the Atomic Theory, you will see exactly how measurements of certain properties helped early scientists to develop theories about the chemical structure of matter on a scale much smaller than they could ever hope to see. You will also learn how these theories, in turn, allow us to make predictions about new materials that humankind has not yet created.

The video below gives you some important tips on how to study chemistry in this class. With practice, you too can learn to think like a chemist, and you may even enjoy it!

Video 1.8.1: How To Study Chemistry.

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1.9: Scientific Problem Solving

How can we use problem solving in our everyday routines?

One day you wake up and realize your clock radio did not turn on to get you out of bed. You are puzzled, so you decide to find out what happened. You list three possible explanations:

1. There was a power failure and your radio cannot turn on.
2. Your little sister turned it off as a joke.
3. You did not set the alarm last night.

Upon investigation, you find that the clock is on, so there is no power failure. Your little sister was spending the night with a friend and could not have turned the alarm off. You notice that the alarm is not set—your forgetfulness made you late. You have used the scientific method to answer a question.

Scientific Problem Solving

Humans have always wondered about the world around them. One of the questions of interest was (and still is): *what is this world made of?* Chemistry has been defined in various ways as the study of matter. What matter consists of has been a source of debate over the centuries. One of the key areas for this debate in the Western world was Greek philosophy.

The basic approach of the Greek philosophers was to discuss and debate the questions they had about the world. There was no gathering of information to speak of, just talking. As a result, several ideas about matter were put forth, but never resolved. The first philosopher to carry out the gathering of data was Aristotle (384-322 B.C.). He recorded many observations on the weather, on plant and animal life and behavior, on physical motions, and a number of other topics. Aristotle could probably be considered the first "real" scientist, because he made systematic observations of nature and tried to understand what he was seeing.

Figure 1.9.1: Aristotle. (Credit: Raphael; Source: http://commons.wikimedia.org/wiki/File:Aristotle_by_Raphael.jpg(opens in new window); License: Public Domain)

Inductive and Deductive Reasoning

Two approaches to logical thinking developed over the centuries. These two methods are **inductive reasoning** and **deductive reasoning**. Inductive reasoning involves getting a collection of specific examples and drawing a general conclusion from them. Deductive reasoning takes a general principle and then draws a specific conclusion from the general concept. Both are used in the development of scientific ideas.

Inductive reasoning first involves the collection of data: "If I add sodium metal to water, I observe a very violent reaction. Every time I repeat the process, I see the same thing happen." A general conclusion is drawn from these observations: the addition of sodium to water results in a violent reaction.

In deductive reasoning, a specific prediction is made based on a general principle. One general principle is that acids turn blue litmus paper red. Using the deductive reasoning process, one might predict: "If I have a bottle of liquid labeled 'acid', I expect the litmus paper to turn red when I immerse it in the liquid."

The Idea of the Experiment

Inductive reasoning is at the heart of what is now called the "**scientific method**." In European culture, this approach was developed mainly by Francis Bacon (1561-1626), a British scholar. He advocated the use of inductive reasoning in every area of life, not just science. The scientific method, as developed by Bacon and others, involves several steps:

1. Ask a question - identify the problem to be considered.
2. Make observations - gather data that pertains to the question.
3. Propose an explanation (a hypothesis) for the observations.
4. Make new observations to test the hypothesis further.

Figure 1.9.2: Sir Francis Bacon. (Credit: Paul van Somer; Source: http://commons.wikimedia.org/wiki/File:Sir_Francis_Bacon.jpg(opens in new window); License: Public Domain)

Note that this should not be considered a "cookbook" for scientific research. Scientists do not sit down with their daily "to do" list and write down these steps. The steps may not necessarily be followed in order. But this does provide a general idea of how scientific research is usually done.

When a hypothesis is confirmed repeatedly, it eventually becomes a theory—a general principle that is offered to explain natural phenomena. Note a key word—*explain*, or *explanation*. A theory offers a description of why something happens. A law, on the other hand, is a statement that is always true, but offers no explanation as to why. The law of gravity says a rock will fall when dropped, but does not explain why (gravitational theory is very complex and incomplete at present). The kinetic molecular theory of gases, on the other hand, states what happens when a gas is heated in a closed container (the pressure increases), but also explains why (the motions of the gas molecules are increased due to the change in temperature). Theories do not get "promoted" to laws, because laws do not answer the "why" question.

Summary

- The early Greek philosophers spent their time talking about nature, but did little or no actual exploration or investigation.
- Inductive reasoning - to develop a general conclusion from a collection of observations.
- Deductive reasoning - to make a specific statement based on a general principle.
- Scientific method - a process of observation, developing a hypothesis, and testing that hypothesis.

Review

1. What was the basic shortcoming of the Greek philosophers approach to studying the material world?
2. How did Aristotle improve the approach?
3. Define “inductive reasoning” and give an example.
4. Define “deductive reasoning” and give an example.
5. What is the difference between a hypothesis and a theory?
6. What is the difference between a theory and a law?

CHAPTER OVERVIEW

2: Measurements

2.1: The Basic Units of Measurement

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

How long is a yard? It depends on whom you ask and when you asked the question. Today we have a standard definition of the yard, which you can see marked on every football field. If you move the ball ten yards, you get a first down and it does not matter whether you are playing in Los Angeles, Dallas, or Green Bay. But at one time that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. Of course, the problem there is simple: new king, new distance (and then you have to re-mark all of those football fields).

Figure 2.1.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 2.1.1.

Table 2.1.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Unfortunately, the Metric System is Not Ubiquitous

The map below shows the adoption of the SI units in countries around the world. The United States has legally adopted the metric system for measurements, but does not use it in everyday practice. Great Britain and much of Canada use a combination of metric and imperial units.

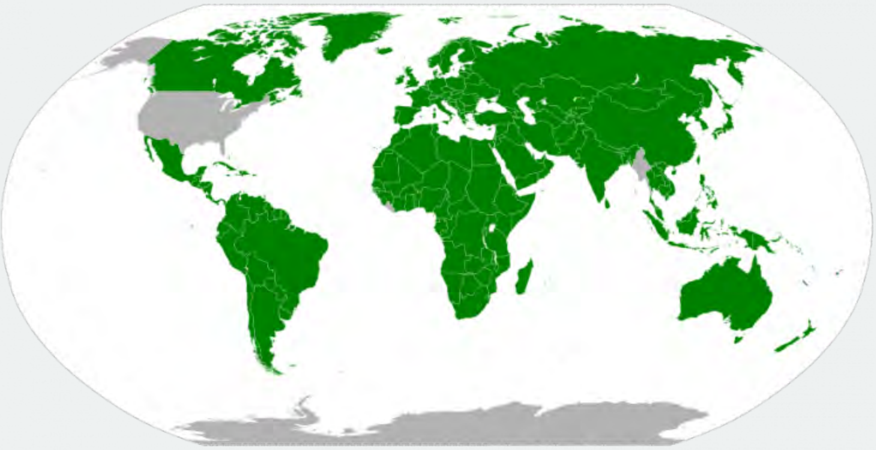


Figure $2.1.1$: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great". Table $2.1.2$ lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table $2.1.2$: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\text{Gm}) = 10^9\ \text{m}$
mega	M	1,000,000	1 megameter $(\text{Mm}) = 10^6\ \text{m}$
kilo	k	1,000	1 kilometer $(\text{km}) = 1{,}000\ \text{m}$
hecto	h	100	1 hectometer $(\text{hm}) = 100\ \text{m}$
deka	da	10	1 dekameter $(\text{dam}) = 10\ \text{m}$
		1	1 meter (m)
deci	d	1/10	1 decimeter $(\text{dm}) = 0.1\ \text{m}$
centi	c	1/100	1 centimeter $(\text{cm}) = 0.01\ \text{m}$
milli	m	1/1,000	1 millimeter $(\text{mm}) = 0.001\ \text{m}$
micro	μ	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6}\ \text{m}$
nano	n	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9}\ \text{m}$
pico	p	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12}\ \text{m}$

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. We use "m" for meter and not "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So we would write 3.5 milliliters as $3.5\ \text{mL}$.

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs $6.5\ \text{kg}$, this is easier than saying it weighs $6500\ \text{g}$ or $0.65\ \text{dag}$. All three are correct, but the kg units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

✓ Example 2.1.1: Unit Abbreviations

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kiloliter
b. microsecond
c. decimeter
d. nanogram

Solutions

Solutions to Example 2.5.1

	Explanation	Answer
a	The prefix kilo means “1,000 ×,” so 1 kL equals 1,000 L.	kL
b	The prefix micro implies 1/1,000,000th of a unit, so 1 µs equals 0.000001 s.	µs
c	The prefix deci means 1/10th, so 1 dm equals 0.1 m.	dm
d	The prefix nano means 1/1000000000, so a nanogram is equal to 0.000000001 g.	ng

? Exercise 2.1.1

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kilometer
b. milligram
c. nanosecond
d. centiliter

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

2.2: Taking Measurements

Learning Objectives

- Express quantities properly, using a number and a unit.

A coffee maker's instructions tell you to fill the coffee pot with 4 cups of water and to use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure $2.2.1$); the nurse is also measuring.

Figure $2.2.1$: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GFDL; Pia von Lützau).

Chemists measure the properties of matter and express these measurements as quantities. A quantity is an amount of something and consists of a number and a unit. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5. If you ask a friend how far they walk from home to school, and the friend answers "12" without specifying a unit, you do not know whether your friend walks 12 kilometers, 12 miles, 12 furlongs, or 12 yards. *Both a number and a unit must be included to express a quantity properly.*

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. The next two sections examine the rules for expressing numbers.

✓ Example $2.2.1$

Identify the number and the unit in each quantity.

a. one dozen eggs
b. 2.54 centimeters
c. a box of pencils
d. 88 meters per second

Solution

a. The number is one, and the unit is a dozen eggs.
b. The number is 2.54, and the unit is centimeter.
c. The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.
d. The number is 88, and the unit is meters per second. Note that in this case the unit is actually a combination of two units: meters and seconds.

Key Take Away

- Identify a quantity properly with a number and a unit.

2.2.1: Measurements in the Laboratory (Experiment)-Home Version

Objectives

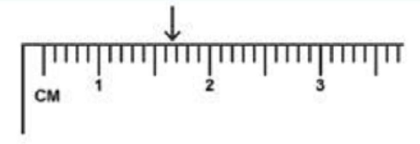
- To use standard laboratory measurement devices to measure length, volume and mass amounts.
- To use these measurements to determine the areas of shapes and volumes
- To determine the density of water.
- To determine the density of a solid and use this to determine further quantities.
- To determine the density of aluminum (applying the technique of water displacement) and use that value to determine the thickness of a piece of aluminum foil.

Chemistry is the study of matter. Our understanding of chemical processes thus depends on our ability to acquire accurate information about matter. Often, this information is quantitative, in the form of *measurements*. In this lab, you will be introduced to some common measuring devices, and learn how to use them to obtain correct measurements, each with correct precision. A metric ruler will be used to measure length in centimeters (cm).

All measuring devices are subject to error, making it impossible to obtain exact measurements. Students will record all the digits of the measurement using the markings that we know exactly and one further digit that we estimate and call uncertain. The *uncertain digit* is our best estimate using the smallest unit of measurement given and estimating between two of these values. These digits are collectively referred to as significant figures. Note, the electronic balance is designed to register these values and the student should only record the value displayed.

When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true, accepted value. Precision refers to how close repeated measurements (using the same device) are to each other.

Example 2.1.1 : Measuring length



Here the “ruler” markings are every 0.1-centimeter. The correct reading is 1.67 cm. The first 2 digits **1.6**7 are known exactly. The last digit 1.6$\underline{7}$ is uncertain. You may have instead estimated it as 1.68 cm.

The measuring devices used in this lab may have different scale graduations than the ones shown Precision is basically how many significant figures you have in your measurement. To find the precision, you basically take the smallest unit on your measuring device, and add a decimal place (the uncertain digit).

Note

In general, the more decimal places provided by a device, the more precise the measurement will be.

Measurements obtained in lab will often be used in subsequent calculations to obtain other values of interest. Thus, it is important to consider the number of significant figures that should be recorded for such calculated values. If multiplying or dividing measured values, the result should be reported with the lowest number of *significant figures* used in the calculation. If adding or subtracting measured values, the result should be reported with the lowest number of *decimal places* used in the calculation.

Example 2.1.2 : Significant Figures in Calculated Values

(a) A student runs 18.752 meters in 54.2 seconds. Calculate his velocity (or speed).

$$velocity = \frac{distance}{time} \tag{2.2.1.1}$$

$$= \frac{18.752m}{54.2s} \tag{2.2.1.2}$$

$$= 0.345978m/s \text{ from calculator} \tag{2.2.1.3}$$

$$= 0.346m/s \text{ to 3 significant figures} \tag{2.2.1.4}$$

(b) The mass of a glass is measured to be 12.456 grams. If 10.33 grams of water are added to this glass, what is the total combined mass?

$$\text{total mass} = 12.456g + 10.33g \tag{2.2.1.5}$$

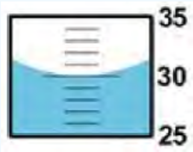
$$= 22.786g \text{ from calculator} \tag{2.2.1.6}$$

$$= 22.79g \text{ to 2 decimal places} \tag{2.2.1.7}$$

In this lab, students will also determine the density of water as well as aluminum. Volume is the amount of space occupied by matter. An extensive property is one that is dependent on the amount of matter present. Volume is an extensive property.

The *volume* of a liquid can be directly measured with specialized glassware, typically in units of milliliters (mL) or liters (L). In this lab, a beaker, two graduated cylinders and a burette will be used to measure liquid volumes, and their precision will be compared. Note that when measuring liquid volumes, it is important to read the graduated scale from the lowest point of the curved surface of the liquid, known as the liquid *meniscus*.

✓ Example 2.1.3 : Measuring the Volume of a liquid



Here, the graduated cylinder markings are every 1-milliliter. When read from the lowest point of the meniscus, the correct volume reading is 30.0 mL. The first 2 digits **30**.0 are known exactly. The last digit 30.**0** is uncertain. Even though it is a zero, it is significant and must be recorded.

The volume of a solid must be measured indirectly based on its shape. For regularly shaped solids, such as a cube, sphere, cylinder, or cone, the volume can be calculated from its measured dimensions (length, width, height, diameter) by using an appropriate equation.

Formulas for Calculating Volumes of Regularly Shaped Solids:

$$\text{Volume of a cube} = l \times w \times h \tag{2.2.1.8}$$

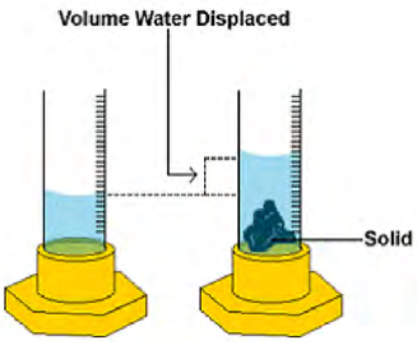
$$\text{Volume of a sphere} = \frac{4}{3}\pi r^3 \tag{2.2.1.9}$$

(where r = radius = 1⁄2 the diameter)

$$\text{Volume of a cylinder} = \pi r^2 h \tag{2.2.1.10}$$

For irregularly shaped solids, the volume can be indirectly determined via the volume of water (or any other liquid) that the solid displaces when it is immersed in the water (*Archimedes Principle*). The units for solid volumes are typically cubic centimeters (cm^3) or cubic meters (m^3). Note that 1 mL = 1 cm3.

Measuring the Volume of an Irregularly Shaped Solid



The volume water displaced is equal to the difference between the final volume and the initial volume , or:

$$V = V_f - V_i \tag{2.2.1.11}$$

where the volume water displaced is equal to the volume of solid.

Density is defined as the mass per unit volume of a substance. Density is a physical property of matter. Physical properties can be measured without changing the chemical identity of the substance. Since pure substances have unique density values, measuring the density of a substance can help identify that substance. Density is also an intensive property. An intensive property is one that is independent of the amount of matter present. For example, the density of a gold coin and a gold statue are the same, even though the gold statue consists of the greater quantity of gold. Density is determined by dividing the mass of a substance by its volume:

$$density = \frac{mass}{volume} \tag{2.2.1.12}$$

Density is commonly expressed in units of g/cm^3 for solids, g/mL for liquids, and g/L for gases.

Procedure

Materials and Equipment

Metric ruler, shape sheet (find a rectangle and circle available at home, say a notebook or circular filter paper), 250-mL Erlenmeyer flask, 100-mL beaker, sugar, 400-mL beaker, spoon, burette (instead of burette, use a long graduated pipette with a bulb), 10-mL and 100-mL graduated cylinders, aluminum pellets/bar, aluminum foil, electronic balance, water (you do not need distilled water, tap water is just fine).

Safety

Be careful when adding the aluminum to your graduated cylinder, as the glass could break. *Personal protective equipment (PPE)* needed: lab coat, safety goggles, closed-toe shoes

Part A: Measuring the Dimensions of Regular Geometric Shapes

1. Find a ruler and “shape sheet” (use any rectangular or circular shaped flat objects like notebook or filter paper) . Measure the dimensions of the two geometric shapes: length and width of the rectangle, and the diameter of the circle. Record these values on your lab report
2. Use your measurements to calculate the area of each shape:

- Area of a rectangle: $A = l \times w$
- Area of a circle: $A = \pi r^2$ (r = radius = 1⁄2 the diameter)

Part B: Volumes of Liquids and Solids

Volumes of Liquids

1. Find a burette (use a long graduated pipette with a bulb instead), 10-mL graduated cylinder, 100-mL graduated cylinder and 100-mL beaker, each filled with a certain quantity of water. Measure the volume of water in each. Remember to read the volume **at the bottom of the meniscus.** It is useful to hold a piece of white paper behind the burette/cylinder/beaker to make it clearer.

Volume of a Regularly Shaped Solid

2. Find a wooden block or cylinder and ruler from your home.
3. Measure the dimensions of the block. If it is a cube or a rectangular box, measure its length, width and height. If it is a cylinder or cone, measure its height and the diameter of its circular base.

Part C: The Density of Water

1. Using the electronic balance determine the mass of a clean, dry, 100-mL graduated cylinder.
2. Pour 40-50 mL of **distilled water** into the graduated cylinder and weigh. Make sure that the outside of the graduated cylinder is dry before placing it on the electronic balance.
3. Measure the liquid volume in the cylinder
4. Use the mass and volume to calculate the density of water.

Part D: The Density of Aluminum and the Thickness of Foil

Density of Aluminum

1. Using the electronic balance to determine the mass of a clean, dry, small beaker.
2. Obtain an aluminum pellet/bar. Transfer pellet/bar to the beaker weighed in the previous step, and measure the mass of the beaker and pellet/bar together.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
4. Carefully add all the aluminum pellet/bar to the water, making sure not to lose any water to splashing. Also make sure that the pellet/bar are all completely immersed in the water. Measure the new volume of the water plus the pellet/bar.
5. When finished, retrieve and dry the aluminum pellet/bar.
6. *Analysis:* Use your measured mass and volume (obtained via water displacement) of the aluminum pellet/bar to calculate the density of aluminum.

The Thickness of Aluminum Foil

7. Take a rectangular piece of aluminum foil and ruler. Use the ruler to measure the length and width of the piece of foil.
8. Fold the foil up into a small square and measure its mass using the electronic balance
9. *Analysis:* Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Lab Report: Measurements in the Laboratory

Part A: Measuring the Dimensions of Regular Geometric Shapes

Experimental Data

Shape	Dimensions	Precision	Measurement	# Significant Figures
Rectangle	Length			
	Width			
Circle	Diameter			

Data Analysis

1. Perform the conversions indicated. Show your work, and report your answers in scientific notation.

- Convert the measured rectangle length to hm.
- Convert the measured circle diameter to nm.

2. Calculate the areas of your rectangle and circle in cm^2. Show your work, and report your answers to the correct number of significant figures.

- Area of rectangle
- Area of circle

Part B: The Volumes of Liquids and Solids

Table 1: The Volume of Liquid Water

Measuring Device	Precision	Volume Measurement	# Significant Figures
Burette			
Beaker			
100-mL Graduated Cylinder			
10-mL Graduated Cylinder			

Table 2: The Volume of a Regular Solid, shaped as a

Dimensions Measured	Measurement	# Significant Figures

Data Analysis

Use your measured block dimensions (in Table 2) to calculate the block volume, in cm^3. Show your work, and report your answer to the correct number of significant figures.

Part C: The Density of Water

Table 1: The Density of Water

Mass of Empty, Dry Graduated Cylinder	
Mass of Graduated Cylinder + Water	
Mass of Water	
Volume of the Water in Graduated Cylinder	

Calculate the density of water, in g/mL. Show your work, and report your answer to the correct number of significant figures.

Part D: The Density of Aluminum and the Thickness of Foil

Experimental Data

Table 1: The Density of Aluminum

Mass of empty beaker	
Mass of beaker and pellet/bar	
Mass of pellet/bar	
Initial volume of water in cylinder	
Final volume of water and pellet/bar	
Volume of pellet	

Table 2: The Thickness of Aluminum Foil

Mass of foil	
Length of foil	
Width of foil	

Data Analysis

1. Use your measured mass and volume of the aluminum pellet (in Table 1) to calculate the density of aluminum, in g/cm^3. Show your work, and report your answer to the correct number of significant figures.
2. Use your measurements for the aluminum foil (in Table 2) along with the true density of aluminum ($\boxed{\_D\_\{Al\}}$ = 2.70 g/cm^3) to calculate the foil thickness, in cm. Consider the foil to be a very flat rectangular box, where Volume of foil = V = length × width × height (thickness). Show your work, and report your answer in scientific notation.

Contributors and Attributions

The experiment is from Chem 9 Experiments. Measurements in the Laboratory (Experiment) is shared under a CC BY-NC license and was authored, remixed, and/or curated by Santa Monica College. Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead, has modified this to fit to a homeschool environment with less hazardous and cost effective lab supplies.

2.3: Scientific Notation - Writing Large and Small Numbers

Learning Objectives

- Express a large number or a small number in scientific notation.
- Carry out arithmetical operations and express the final answer in scientific notation

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. In scientific notation, these numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 ($1 \leq N < 10$), and n is a positive or negative integer ($10^0 = 1$). The number 10 is called the base because it is this number that is raised to the power n. Although a base number may have values other than 10, the base number in scientific notation is always 10.

A simple way to convert numbers to scientific notation is to move the decimal point as many places to the left or right as needed to give a number from 1 to 10 (N). The magnitude of n is then determined as follows:

- If the decimal point is moved to the left n places, n is positive.
- If the decimal point is moved to the right n places, n is negative.

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example 2.3.1.

✓ Example 2.3.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

a. 637.8
b. 0.0479
c. 7.86
d. 12,378
e. 0.00032
f. 61.06700
g. 2002.080
h. 0.01020

Solution

Solutions to Example 2.2.1

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, n = 2.	6.378×10^2
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, n = −2.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^0 = 1$.)	7.86×10^0

	Explanation	Answer
d	Because the decimal point was moved four places to the left, n = 4.	1.2378×10^4
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^1
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^3
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example $2.3.2$ illustrates how to do this.

Example $2.3.2$: Expressing Sums and Differences in Scientific Notation

Carry out the appropriate operation and then express the answer in scientific notation.

a. $(1.36 \times 10^2) + (4.73 \times 10^3)$
b. $(6.923 \times 10^{-3}) - (8.756 \times 10^{-4})$

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) = (1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) = (0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87×10^3
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) = (6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) = (69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.0474×10^{-3}

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example $2.3.3$.

✓ Example 2.3.3: Expressing Products and Quotients in Scientific Notation

Perform the appropriate operation and express your answer in scientific notation.

a. $(6.022 \times 10^{23})(6.42 \times 10^{-2})$

b. $\dfrac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$

c. $\dfrac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

Solution to Example 2.2.3

	Explanation	Answer
a	In multiplication, we add the exponents: $(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	

3.87×10^{22} **b**

In division, we subtract the exponents:

$$\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}} = \frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^{4}$$

In converting 0.183×10^{4} to scientific notation, n has become more negative by 1 because the value of N has increased.

1.83×10^{3} **c**

This problem has both multiplication and division:

$$\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{(8.52 \times 10^{-2})} = \frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$$

4.7×10^{-31}

2.4: Significant Figures - Writing Numbers to Reflect Precision

Learning Objectives

- Identify the number of significant figures in a reported value.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length at 2.553 cm as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

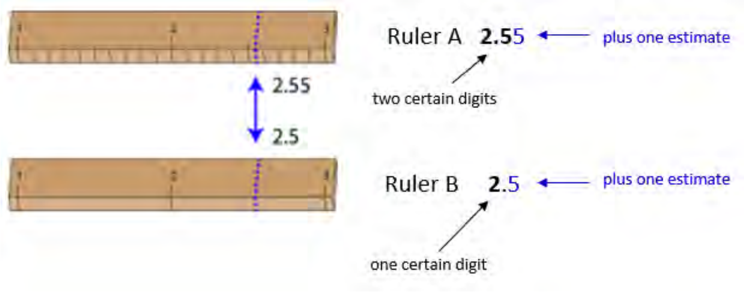


Figure 2.4.1: Measurement with two different rulers.
Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 2.4.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

With either ruler, it is clear that the length of the object is between 2 and 3 cm. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as 2.5 cm. However, another person may judge that the measurement is 2.4 cm or perhaps 2.6 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

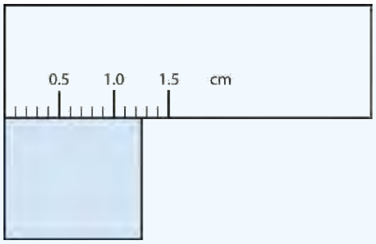
The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

Example 2.4.1: Reporting Measurements to the Proper Number of Significant Figures

Use each diagram to report a measurement to the proper number of significant figures.

a.

b.



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

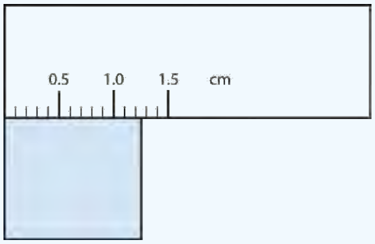
Solutions

Solutions to Example 2.3.1

	Explanation	Answer
a.	The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it’s at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. The symbol psi stands for “pounds per square inch” and is a unit of pressure, like air in a tire. The measurement is reported to three significant figures.	4.33 psi
b.	The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle’s width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. The measurement is reported to three significant figures.	1.25 cm

? Exercise 2.4.1

What would be the reported width of this rectangle?



Answer

1.25 cm

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 2.4.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous. ∘ 1.4×10^3 has two significant figures. ∘ 1.40×10^3 three significant figures. ∘ 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000…). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

✓ Example 2.4.2

Give the number of significant figures in each. Identify the rule for each.

a. 5.87
b. 0.031
c. 52.90
d. 00.2001
e. 500
f. 6 atoms

Solution

Solution to Example 2.3.2

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d	The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e	The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f	The 6 is a counting number. A counting number is an exact number.	6, infinite

? Exercise 2.4.2

Give the number of significant figures in each.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people
e. 3,800 g

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

Answer e

Ambiguous, could be two, three or four significant figures.

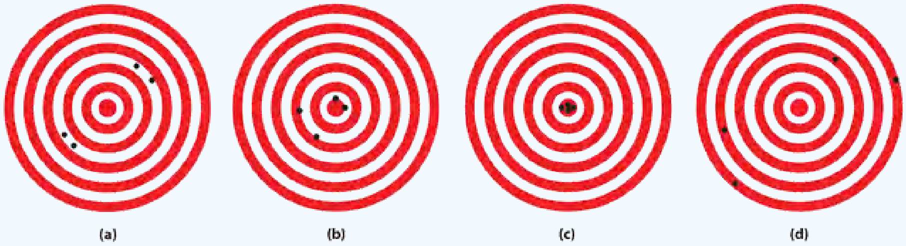
Accuracy and Precision

Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrates the concepts of accuracy and precision.

Video 2.4.1: Difference between precision and accuracy.

✓ Example 2.4.3

The following archery targets show marks that represent the results of four sets of measurements.



Which target shows

a. a precise, but inaccurate set of measurements?
b. a set of measurements that is both precise and accurate?
c. a set of measurements that is neither precise nor accurate?

Solution

a. Set a is precise, but inaccurate.
b. Set c is both precise and accurate.
c. Set d is neither precise nor accurate.

Summary

Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool. Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

2.5: Significant Figures in Calculations

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement $207.518\ \mathrm{m}$ Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table 2.5.1.

Table 2.5.1: Rounding examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Notice that the more rounding that is done, the less reliable the figure is. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

It is important to be aware of significant figures when you are mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238… to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers that have only three significant figures each? When performing mathematical operations, there are two rules for limiting the number of significant figures in an answer—one rule is for addition and subtraction, and one rule is for multiplication and division.

> *In operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the **least precise** operation. An answer is no more precise than the least precise number used to get the answer.*

Multiplication and Division

For multiplication or division, the rule is to count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

$$\underbrace{38.65}_{\text{4 sig figs}} \times \underbrace{105.93}_{\text{5 sig figs}} = \underbrace{4{,}094.1945}_{\text{reduce to 4 sig figs}}$$

The final answer, limited to four significant figures, is 4,094. The first digit dropped is 1, so we do not round up.

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 450 has two significant figures and would be written in scientific notation as 4.5×10^2, whereas 450.0 has four significant figures and would be written as 4.500×10^2. In scientific notation, all significant figures are listed explicitly.

Example 2.5.1

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 23.096 × 90.300
b. 125 × 9.000

Solution

a

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	2.0856×10^3

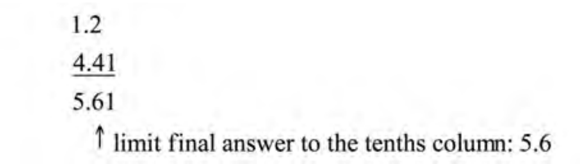
b

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator gives 1,125 as the answer, but we limit it to three significant figures.	1.13×10^3

Addition and Subtraction

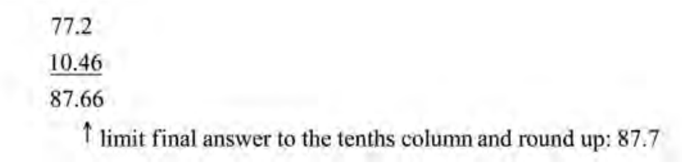
How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.71, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.



1.2
4.41
5.61
↑ limit final answer to the tenths column: 5.6

We drop the last digit—the 1—because it is not significant to the final answer.

The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater, and rounded down if the first dropped digit is less than 5.



77.2
10.46
87.66
↑ limit final answer to the tenths column and round up: 87.7

Example 2.5.2

a. 13.77 + 908.226
b. 1,027 + 611 + 363.06

Solution

a

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place, we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$

b

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2{,}001.06 = 2.001 \times 10^3$

Exercise 2.5.2

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 217 ÷ 903
b. 13.77 + 908.226 + 515
c. 255.0 − 99
d. 0.00666 × 321

Answer a:

$$0.240 = 2.40 \times 10^{-1}$$

Answer b:

$$1,437 = 1.437 \times 10^3$$

Answer c:

$$156 = 1.56 \times 10^2$$

Answer d:

$$2.14 = 2.14 \times 10^0$$

Remember that calculators do not understand significant figures. *You* are the one who must apply the rules of significant figures to a result from your calculator.

Calculations Involving Multiplication/Division and Addition/Subtraction

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain at least one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.

Video 2.5.1: Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).

Video 2.5.2: https://www.youtube.com/watch?v=__csP0NtlGI

In the worked examples in this text, we will often show the results of intermediate steps in a calculation. In doing so, we will show the results to only the correct number of significant figures allowed for that step, in effect treating each step as a separate calculation. This procedure is intended to reinforce the rules for determining the number of significant figures, but in some cases it may give a final answer that differs in the last digit from that obtained using a calculator, where all digits are carried through to the last step.

✓ Example 2.5.3

a. 2(1.008 g) + 15.99 g
b. 137.3 s + 2(35.45 s)
c. $\frac{118.7g}{2} - 35.5g$

Solution

a.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for multiplication first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer

2(1.008 g) + 15.99 g = Perform multiplication first. 2 (1.008 g *4 sig figs*) = **2.01$\underline{6}$ g** *4 sig figs* The number with the least number of significant figures is 1.008 g; *the number 2 is an exact number and therefore has an infinite number of significant figures*. **Then, perform the addition.** **2.01$\underline{6}$ g** *thousandths place* + 15.9$\underline{9}$ g *hundredths place (least precise)* = 18.006 g **Round the final answer.** Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)

b.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for multiplication first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
137.3 s + 2(35.45 s) = **Perform multiplication first.** 2(35.45 s *4 sig figs*) = **70.90 s** *4 sig figs* The number with the least number of significant figures is 35.45; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Then, perform the addition.** 137.3 s *tenths place (least precise)* + **70.90 s** *hundredths place* = 208.20 s **Round the final answer.** Round the final answer to the tenths place based on 137.3 s.	208.2 s

c.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for division first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	**Answer**
$\frac{118.7g}{2} - 35.5g =$ **Perform division first.** $\frac{118.7g}{2}$ *4 sig figs* = **59.35 g** *4 sig figs* The number with the least number of significant figures is 118.7 g; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Perform subtraction next.** **59.35 g** *hundredths place* − 35.5 g *tenths place (least precise)* = 23.85 g **Round the final answer.** Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

? Exercise 2.5.3

Complete the calculations and report your answers using the correct number of significant figures.

a. 5(1.008s) - 10.66 s
b. 99.0 cm+ 2(5.56 cm)

Answer a

-5.62 s

Answer b

110.2 cm

Summary

- Rounding
  - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1 (e.g. 2.9699 rounded to three significant figures is 2.97).
  - If the number to be dropped is less than 5, there is no change (e.g. 4.00443 rounded to four significant figures is 4.004).
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the number with the fewest significant figures.
- The rule in addition and subtraction is that the answer is given the same number of decimal places as the term with the fewest decimal places.

2.6: Problem Solving and Unit Conversions

Learning Objectives

- To convert a value reported in one unit to a corresponding value in a different unit using conversion factors.

During your studies of chemistry (and physics also), you will note that mathematical equations are used in many different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Algebra skills become very important here!

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, or a more complex one can convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with (Table $2.6.1$).

Table $2.6.1$: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table $2.6.1$. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions in the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds $=$ 1 minute).

If you learned the SI units and prefixes described, then you know that 1 cm is 1/100th of a meter.

$$1 \text{ cm} = \frac{1}{100} \text{ m} = 10^{-2} \text{m}$$

or

$$100 \text{ cm} = 1 \text{ m}$$

Suppose we divide both sides of the equation by 1m (both the number *and* the unit):

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1 \text{ m}}{1 \text{ m}}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1000 \text{ mm}}{1 \text{ m}} = \frac{1 \times 10^6 \mu\text{m}}{1 \text{ m}} = 1$$

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units.

Performing Dimensional Analysis

Dimensional analysis is amongst the most valuable tools that physical scientists use. Simply put, it is the conversion between an amount in one unit to the corresponding amount in a desired unit using various conversion factors. This is valuable because certain measurements are more accurate or easier to find than others. The use of units in a calculation to ensure that we obtain the final proper units is called *dimensional analysis*.

Here is a simple example. How many centimeters are there in 3.55 m? Perhaps you can determine the answer in your head. If there are 100 cm in every meter, then 3.55 m equals 355 cm. To solve the problem more formally with a conversion factor, we first write the quantity we are given, 3.55 m. Then we multiply this quantity by a conversion factor, which is the same as multiplying it by 1. We can write 1 as $\frac{100\text{ cm}}{1\text{ m}}$ and multiply:

$$3.55\text{ m}\times\frac{100\text{ cm}}{1\text{ m}}$$

The 3.55 m can be thought of as a fraction with a 1 in the denominator. Because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out:

$$\frac{3.55\ \cancel{\text{m}}}{1}\times\frac{100\text{ cm}}{1\ \cancel{\text{m}}}$$

The final step is to perform the calculation that remains once the units have been canceled:

$$\frac{3.55}{1}\times\frac{100\text{ cm}}{1}=355\text{ cm}$$

In the final answer, we omit the 1 in the denominator. Thus, by a more formal procedure, we find that 3.55 m equals 355 cm. A generalized description of this process is as follows:

quantity (in old units) × conversion factor = quantity (in new units)

You may be wondering why we use a seemingly complicated procedure for a straightforward conversion. In later studies, the conversion problems you encounter *will not always be so simple*. If you master the technique of applying conversion factors, you will be able to solve a large variety of problems.

In the previous example, we used the fraction $\frac{100\text{ cm}}{1\text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1\text{ m}}{100\text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

$$3.55\text{ m}\times\frac{1\text{ m}}{100\text{ cm}}=0.0355\frac{\text{m}^2}{\text{cm}}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure $2.6.1$ shows a concept map for constructing a proper conversion.

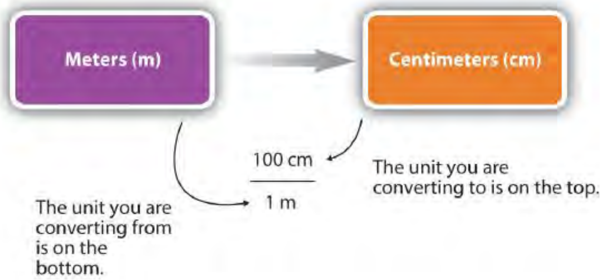


Figure $2.6.1$: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Converting meters to centimeters: the unit you are converting is on the bottom of the fraction, and the unit to convert to is on top

General Steps in Performing Dimensional Analysis

1. Identify the "**given**" information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to "**find**"? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

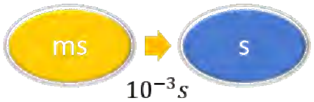
Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

- Numbers in conversion factors based on prefix changes, such as kilograms to grams, are *not* considered in the determination of significant figures in a calculation because the numbers in such conversion factors are exact.
- Exact numbers are defined or counted numbers, not measured numbers, and can be considered as having an infinite number of significant figures. (In other words, 1 kg is exactly 1,000 g, by the definition of kilo-.)
- Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact.
- In contrast, conversion factors that come from measurements (such as density, as we will see shortly) or that are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

Example 2.6.1

Steps for Problem Solving for Example 2.6.1 and 2.6.2

	Example 2.6.1	**Example 2.6.2**
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\,mL = 10^{-3}\,L$	$1\,ms = 10^{-3}\,s$
Prepare a concept map and use the proper conversion factor.	 $\frac{1 mL}{10^{-3} L}$	 $\frac{10^{-3} s}{1 ms}$
Cancel units and calculate.	$4.7\ \cancel{\text{L}} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{\text{L}}} = 4{,}700\ \text{mL}$ or $4.7\ \cancel{\text{L}} \times \frac{1{,}000\ \text{mL}}{1\ \cancel{\text{L}}} = 4{,}700\ \text{mL}$ or 4.7 x 10^3 2SF, not ambiguous	$18\ \cancel{\text{ms}} \times \frac{10^{-3}\ \text{s}}{1\ \cancel{\text{ms}}} = 0.018\ \text{s}$ or $18\ \cancel{\text{ms}} \times \frac{1\ \text{s}}{1{,}000\ \cancel{\text{ms}}} = 0.018\ \text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

Exercise 2.6.1

Perform each conversion.

a. 101,000 ns to seconds
b. 32.08 kg to grams
c. 1.53 grams to cg

Answer a:

$1.01000x10^{-4}s$

Answer b:

$3.208x10^{4}g$

Answer c:

$1.53x10^{2}cg$

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce the appropriate units.

2.7: Solving Multistep Conversion Problems

Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence. We first convert the given amount in km to the base unit, which is meters. We know that 1,000 m =1 km.

Then we convert meters to mm, remembering that $1\ \mathrm{mm} = 10^{-3}\ \mathrm{m}$.

Concept Map

$$\frac{1000\,m}{1\,km} \qquad \frac{1\,mm}{10^{-3}}$$

Convert kilometers to meters to millimeters: use conversion factors 1000 meters per 1 kilometer and 1 millimeter per 0.001 meter

Calculation

$$\begin{aligned} 54.7\ \cancel{\mathrm{km}} \times \frac{1,000\ \cancel{\mathrm{m}}}{1\ \cancel{\mathrm{km}}} \times \frac{1\ \cancel{\mathrm{mm}}}{10^{-3}\cancel{\mathrm{m}}} &= 54,700,000\ \mathrm{mm} \\ &= 5.47 \times 10^{7}\ \mathrm{mm} \end{aligned}$$

> *In each step, the previous unit is canceled and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.*

✓ Example 2.7.1: Unit Conversion

Convert 58.2 ms to megaseconds in one multi-step calculation.

Solution

Solution for Example 2.7.1

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 58.2 ms Find: Ms
List other known quantities	$1 ms = 10^{-3} s$ $1 Ms = 10^{6} s$
Prepare a concept map.	 $\frac{10^{-3} s}{1\ ms}$ $\frac{1\ Ms}{10^{6}\ s}$ Convert milliseconds to seconds to microseconds: use conversion factors 0.001 second per millisecond and 1 microsecond per 1 million seconds

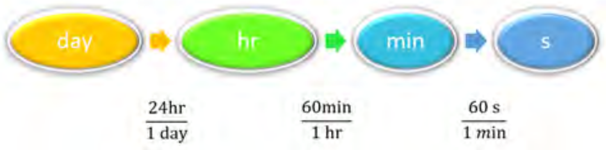
Steps for Problem Solving	Unit Conversion
Calculate.	$$58.2\ \cancel{ms} \times \frac{10^{-3}\ \cancel{s}}{1\ \cancel{ms}} \times \frac{1\ Ms}{1{,}000{,}000\ \cancel{s}} = 0.0000000582\ Ms$$ $$= 5.82 \times 10^{-8}\ Ms$$ Neither conversion factor affects the number of significant figures in the final answer.

Example 2.7.2: Unit Conversion

How many seconds are in a day?

Solution

Solution for Example 2.7.2

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 1 day Find: s
List other known quantities.	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds
Prepare a concept map.	 Convert day to hour to minute to second: use conversion factors 24 hours per day, 60 minutes per hour, and 60 seconds per minute
Calculate.	$$1\ \text{d} \times \frac{24\ \text{hr}}{1\ \text{d}} \times \frac{60\ \text{min}}{1\ \text{hr}} \times \frac{60\ \text{s}}{1\ \text{min}} = 86{,}400\ \text{s}$$

Exercise 2.7.1

Perform each conversion in one multi-step calculation.

a. 43.007 ng to kg
b. 1005 in to ft
c. 12 mi to km

Answer a

$4.3007 \times 10^{-11} kg$

Answer b

$83.75\ ft$

Answer c

$19\ km$

Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school. Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities. Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone’s health.

A woman consulting with a pharmacist. (Public Domain; Rhoda Baer via National Cancer Institute, an agency that is part of the National Institutes of Health.)

Summary

In multi-step conversion problems, the previous unit is canceled for each step and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.

2.8: Numerical Problem-Solving Strategies and the Solution Map

2.9: Density

Learning Objectives

- Define density.
- Use density as a conversion factor.

Density (ρ) is a physical property found by dividing the mass of an object by its volume. Regardless of the sample size, density is always constant. For example, the density of a pure sample of tungsten is always 19.25 grams per cubic centimeter. This means that whether you have one gram or one kilogram of the sample, the density will never vary. The equation, as we already know, is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

or just

$$\rho = \frac{m}{V} \tag{2.9.1}$$

Based on this equation, it's clear that density can, and does, vary from element to element and substance to substance due to differences in the relationship of mass and volume. Pure water, for example, has a density of 0.998 g/cm^3 at 25° C. The average densities of some common substances are in Table $2.9.1$. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table 2.9.1: Densities of Common Substances

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Density can be measured for all substances—solids, liquids and gases. For solids and liquids, density is often reported using the units of g/cm^3. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

✓ Example 2.9.1: Ethyl Alcohol

Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g

Solution

This is a direct application of Equation 2.9.1:

$$\rho = \frac{23.71002\, g}{30.2\, mL} = 0.785\, g/mL$$

? Exercise 2.9.1

a. Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.
b. If you have a 2.130 mL sample of acetic acid with mass 0.002234 kg, what is the density in kg/L?

Answer a

$0.274\, kg/L$

Answer b

$1.049\ kg/L$

Density as a Conversion Factor

Conversion factors can also be constructed for converting between different kinds of units. For example, density can be used to convert between the mass and the volume of a substance. Consider mercury, which is a liquid at room temperature and has a density of 13.6 g/mL. The density tells us that 13.6 g of mercury have a volume of 1 mL. We can write that relationship as follows:

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

$$\frac{13.6\ g}{1\ mL} = 1$$

and

$$\frac{1\ mL}{13.6\ g} = 1$$

Which one do we use? It depends, as usual, on the units we need to cancel and introduce. For example, suppose we want to know the mass of 2.0 mL of mercury. We would use the conversion factor that has milliliters on the bottom (so that the milliliter unit cancels) and grams on top, so that our final answer has a unit of mass:

$$2.0\ \cancel{\text{mL}} \times \frac{13.6\ \text{g}}{1\ \cancel{\text{mL}}} = 27.2\ \text{g} = 27\ \text{g}$$

In the last step, we limit our final answer to two significant figures because the volume quantity has only two significant figures; the 1 in the volume unit is considered an exact number, so it does not affect the number of significant figures. The other conversion factor would be useful if we were given a mass and asked to find volume, as the following example illustrates.

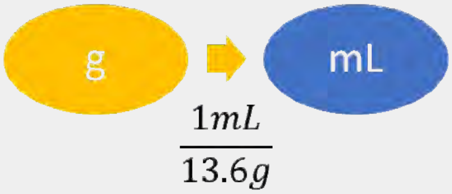
> *Density can be used as a conversion factor between mass and volume.*

✓ Example 2.9.2: Mercury Thermometer Steps for Problem Solving

A mercury thermometer for measuring a patient’s temperature contains 0.750 g of mercury. What is the volume of this mass of mercury?

Solution

Solution to Example 2.9.2

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 0.750 g Find: mL
List other known quantities.	13.6 g/mL (density of mercury)
Prepare a concept map.	
Calculate.	$0.750\ \cancel{\text{g}} \times \frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}} = 0.055147\ldots\ \text{mL} \approx 0.0551\ \text{mL}$ We have limited the final answer to three significant figures.

? Exercise 2.9.2

What is the volume of 100.0 g of air if its density is 1.3 g/L?

Answer

$77\ L$

Summary

- Density is defined as the mass of an object divided by its volume.
- Density can be used as a conversion factor between mass and volume.

Contributions & Attributions

2.E: Measurement and Problem Solving (Exercises)

2.1: Measuring Global Temperatures

2.2: Scientific Notation: Writing Large and Small Numbers

2.3: Significant Figures: Writing Numbers to Reflect Precision

1. Define *significant figures*. Why are they important?
2. Define the different types of zeros found in a number and explain whether or not they are significant.
3. How many significant figures are in each number?
    a. 140
    b. 0.009830
    c. 15,050
    d. 221,560,000
    e. 5.67×10^3
    f. 2.9600×10^{-5}
4. How many significant figures are in each number?
    a. 1.05
    b. 9,500
    c. 0.0004505
    d. 0.00045050
    e. 7.210×10^6
    f. 5.00×10^{-6}
5. Round each number to three significant figures.
    a. 34,705
    b. 34,750
    c. 34,570

2.4: Significant Figures in Calculations

2.5: The Basic Units of Measurement

2.6: Problem Solving and Unit Conversions

2.7: Solving Multi-step Conversion Problems

2.8: Units Raised to a Power

2.9: Density

2.10: Numerical Problem-Solving Strategies and the Solution Map

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CHAPTER OVERVIEW

3: Matter and Energy

3.1: In Your Room

Matter is any substance that has mass and takes up space. Matter includes atoms and anything made up of atoms, but not other energy phenomena or waves such as light or sound. While this simple definition is easily applied, the way people view matter is often broken down into two characteristic length scales: the macroscopic and the microscopic.

Figure 3.1.1: A typical American university and college dormitory room in 2002 (CC BY-SA 3.0; Raul654).

The macroscopic scale is the length scale on which objects or phenomena are large enough to be visible almost practically with the naked eye, without magnifying optical instruments. Everything that one can see, touch, and handle in the dorm room of Figure 3.1.1 is within the macroscopic scale. To describe each of these objects, only a few macroscopic properties are required. However, each of these items can be decomposed into smaller microscopic scale properties.

The microscopic scale is the scale of objects and events smaller than those that can easily be seen by the naked eye, requiring a lens or microscope to see them clearly. All of the everyday objects that we can bump into, touch, or squeeze are ultimately composed of atoms. This ordinary atomic matter is in turn made up of interacting subatomic particles—usually a nucleus of protons and neutrons, and a cloud of orbiting electrons. Because of this, a large number of variables are needed to describe such a system which complicates the characterization.

Matter vs. Mass

Matter should not be confused with mass, as the two are not the same in modern physics. Matter is a physical substance of which systems may be composed, while mass is not a substance, but rather a quantitative property of matter and other substances or systems.

Contributions & Attributions

- Wikipedia

3.2: What is Matter?

Learning Objectives

- Define matter and explain how it is composed of building blocks known as "atoms".

We are all familiar with matter. The definition of **Matter** is anything that has mass and volume (takes up space). For most common objects that we deal with every day, it is fairly simple to demonstrate that they have mass and take up space. You might be able to imagine, however, the difficulty for people several hundred years ago to demonstrate that air had mass and volume. Air (and all other gases) are invisible to the eye, have very small masses compared to equal amounts of solids and liquids, and are quite easy to compress (change volume). Without sensitive equipment, it would have been difficult to convince people that gases are matter. Today, we can measure the mass of a small balloon when it is deflated and then blow it up, tie it off, and measure its mass again to detect the additional mass due to the air inside. The mass of air, under room conditions, that occupies a one quart jar is approximately 0.0002 pounds. This small amount of mass would have been difficult to measure in times before balances were designed to accurately measure very small masses. Later, scientists were able to compress gases into such a small volume that the gases turned into liquids, which made it clear that gases are matter.

Figure 3.2.1: Everything from an ant, to a truck, to the earth, and even the entire galaxy is composed of matter. Images used with permission from Wikipedia (CC_SA-BY-3.0; credit High Contrast).

Even though the universe consists of "things" as wildly different as ants and galaxies, the matter that makes up all of these "things" is composed of a very limited number of building blocks. These building blocks are known as **atoms**, and so far, scientists have discovered or created a grand total of 118 different types of atoms. Scientists have given a name to each different type of atom. A substance that is composed of only one type of atom is called an element. At this point, what should amaze you is that all forms of matter in our universe are made with only 118 different building blocks. In some ways, it's sort of like cooking a gourmet, five-course meal using only three ingredients! How is it possible? To answer that question, you have to understand the ways in which different elements are put together to form matter.

The most important method that nature uses to organize atoms into matter is the formation of **molecules**. Molecules are groups of two or more atoms that have been bonded together. There are millions of different ways to bond atoms together, which means that there are millions of different possible molecules. Each of these molecules has its own set of chemical properties, and it's these properties with which chemists are most concerned. You will learn a lot more about atoms and molecules, including how they were discovered, in a later part of the textbook.

Summary

All matter has mass and occupies space. All physical objects are made of matter. Matter itself is composed of tiny building blocks known as "atoms". There are only 118 different types of atoms known to man. Frequently, atoms are bonded together to form "molecules".

Contributions & Attributions

- Wikipedia

3.3: Classifying Matter According to Its State- Solid, Liquid, and Gas

Learning Objectives

- To describe the solid, liquid and gas phases.

Water can take many forms. At low temperatures (below $0^{o}C$), it is a solid. When at "normal" temperatures (between $0^{o}C$ and $100^{o}C$), it is a liquid. While at temperatures above $100^{o}C$, water is a gas (steam). The state that water is in depends upon the temperature. Each state has its own unique set of physical properties. Matter typically exists in one of three states: **solid**, **liquid**, or **gas**.

Figure 3.3.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure 3.3.2 shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

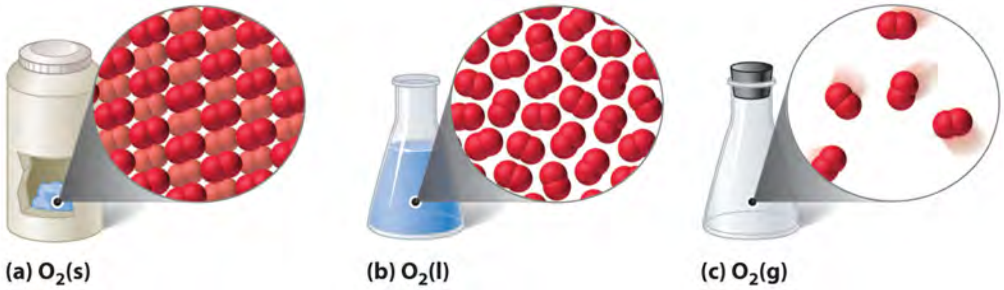

Figure 3.3.2: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

Plasma: A Fourth State of Matter

Technically speaking, a fourth state of matter called plasma exists, but it does not naturally occur on earth, so we will omit it from our study here.

A plasma globe operating in a darkened room. (CC BY-SA 3.0; Chocolateoak).

Solids

In the solid state, the individual particles of a substance are in fixed positions with respect to each other because there is not enough thermal energy to overcome the intermolecular interactions between the particles. As a result, solids have a definite shape and volume. Most solids are hard, but some (like waxes) are relatively soft. Many solids composed of ions can also be quite brittle.

Solids are defined by the following characteristics:

- Definite shape (rigid)
- Definite volume
- Particles vibrate around fixed axes

If we were to cool liquid mercury to its freezing point of -39°C, and under the right pressure conditions, we would notice all of the liquid particles would go into the solid state. Mercury can be solidified when its temperature is brought to its freezing point. However, when returned to room temperature conditions, mercury does not exist in solid state for long, and returns back to its more common liquid form.

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure $3.3.3$. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

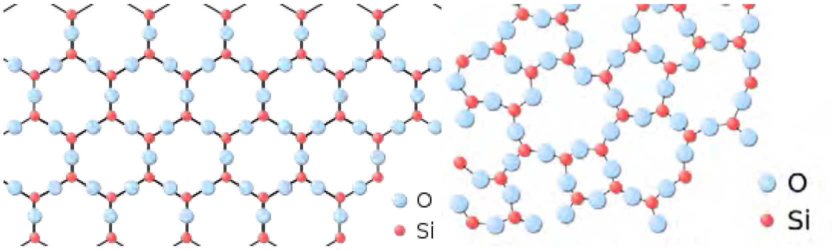

Figure $3.3.3$: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

If the particles of a substance have enough energy to partially overcome intermolecular interactions, then the particles can move about each other while remaining in contact. This describes the liquid state. In a liquid, the particles are still in close contact, so liquids have a definite volume. However, because the particles can move about each other rather freely, a liquid has no definite shape and takes a shape dictated by its container.

Liquids have the following characteristics:

- No definite shape (takes the shape of its container).
- Has definite volume.
- Particles are free to move over each other, but are still attracted to each other.

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension)—a property that all liquids exhibit. Mercury has a relatively high surface tension, which makes it very unique. Here you see mercury in its common liquid form.

Video* $3.3.1$*: Mercury boiling to become a gas.

If we heat liquid mercury to its boiling point of $357^{\circ}\mathrm{C}$ under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. This describes the gas state, which we will consider in more detail elsewhere. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either. The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Gases have the following characteristics:

- No definite shape (takes the shape of its container)
- No definite volume
- Particles move in random motion with little or no attraction to each other
- Highly compressible

Table $3.3.1$: Characteristics of the Three States of Matter

Characteristics	Solids	Liquids	Gases
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

✓ Example $3.3.1$

What state or states of matter does each statement, describe?

a. This state has a definite volume, but no definite shape.
b. This state has no definite volume.
c. This state allows the individual particles to move about while remaining in contact.

Solution

a. This statement describes the liquid state.
b. This statement describes the gas state.
c. This statement describes the liquid state.

Exercise 3.3.1

What state or states of matter does each statement describe?

a. This state has individual particles in a fixed position with regard to each other.
b. This state has individual particles far apart from each other in space.
c. This state has a definite shape.

Answer a:

solid

Answer b:

gas

Answer c:

solid

Summary

- Three states of matter exist—solid, liquid, and gas.
- Solids have a definite shape and volume.
- Liquids have a definite volume, but take the shape of the container.
- Gases have no definite shape or volume.

3.4: Classifying Matter According to Its Composition

Learning Objectives

- Explain the difference between a pure substance and a mixture.
- Explain the difference between an element and a compound.
- Explain the difference between a homogeneous mixture and a heterogeneous mixture.

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 3.4.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

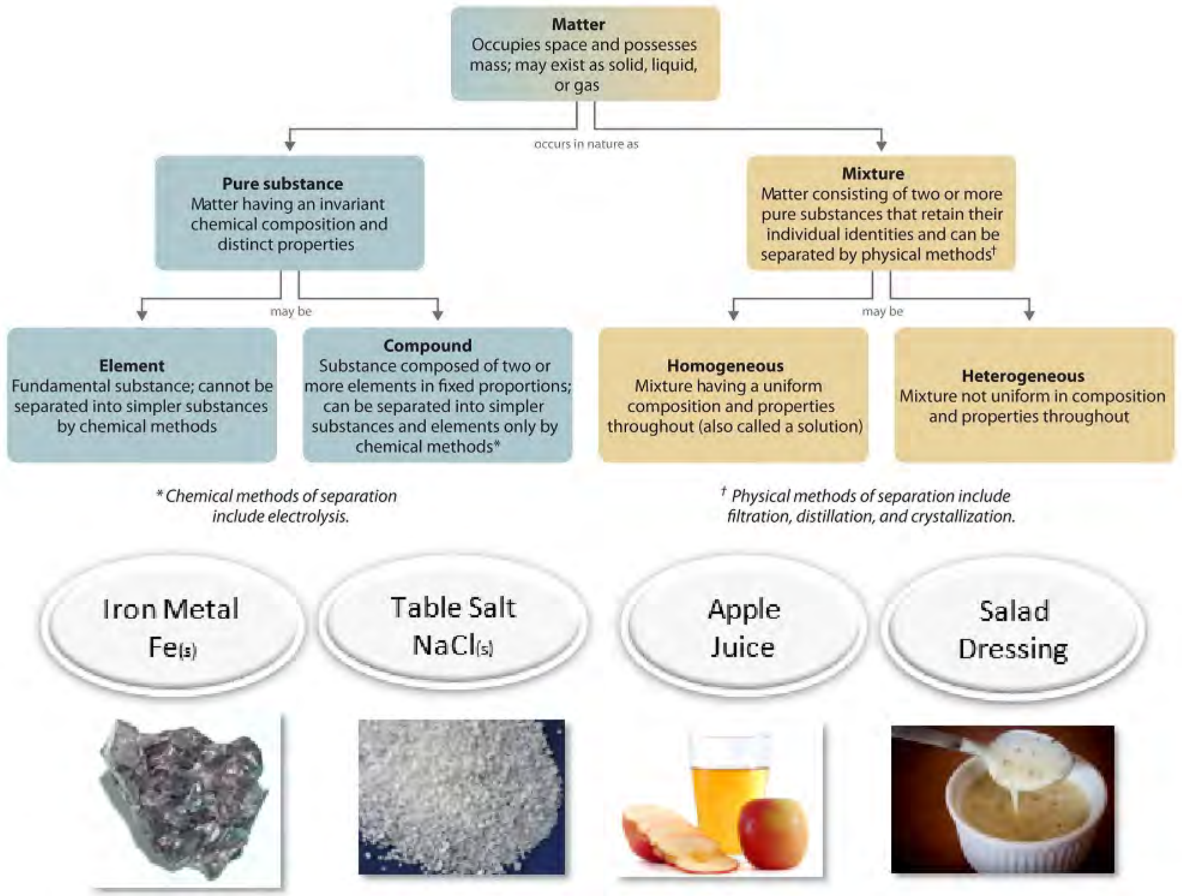

Figure 3.4.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

A **homogeneous** mixture is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. The amount of salt in the salt water can vary from one sample to another. All solutions are considered homogeneous because the dissolved material is present in the same amount throughout the solution.

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

Phase

A phase is any part of a sample that has a uniform composition and properties. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly, but instead form two separate layers. Each of the layers is called a phase.

✓ Example 3.4.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. filtered tea
b. freshly squeezed orange juice
c. a compact disc
d. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
e. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
   B) Because the composition of the solution is uniform throughout, it is ***a homogeneous mixture***.
b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
   B) Because its composition is not uniform throughout, orange juice is ***a heterogeneous mixture***.
c. **A)** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence, a compact disc is not chemically pure.
   B) The regions of different composition indicate that a compact disc is ***a heterogeneous mixture.***
d. **A)** Aluminum oxide is a single, chemically ***pure compound***.
e. **A)** Selenium is one of the known ***elements***.

? Exercise 3.4.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. white wine
b. mercury
c. ranch-style salad dressing
d. table sugar (sucrose)

Answer a:

homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture

Answer d:

compound

✓ Example 3.4.2

How would a chemist categorize each example of matter?

a. saltwater
b. soil
c. water
d. oxygen

Solution

a. Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a homogeneous mixture, or a solution.
b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
c. Water is a substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a substance, is an element.

? Exercise 3.4.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a:

a homogeneous mixture (solution), assuming it is filtered coffee

Answer b:

element

Answer c:

heterogeneous mixture

Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

- Element: a substance that is made up of only one type of atom.
- Compound:a substance that is made up of more than one type of atom bonded together.
- Mixture: a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties.

Contributions & Attributions

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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

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3.5: Differences in Matter- Physical and Chemical Properties

Learning Objectives

To separate physical from chemical properties.

All matter has physical and chemical properties. Physical properties are characteristics that scientists can measure without changing the composition of the sample under study, such as mass, color, and volume (the amount of space occupied by a sample). Chemical properties describe the characteristic ability of a substance to react to form new substances; they include its flammability and susceptibility to corrosion. All samples of a pure substance have the same chemical and physical properties. For example, pure copper is always a reddish-brown solid (a physical property) and always dissolves in dilute nitric acid to produce a blue solution and a brown gas (a chemical property).

Physical Property

A *physical property* is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets, a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C. The average densities of some common substances are in Table $3.5.1$. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Table $3.5.1$: Densities of Common Substances

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Hardness helps determine how an element (especially a metal) might be used. Many elements are fairly soft (silver and gold, for example) while others (such as titanium, tungsten, and chromium) are much harder. Carbon is an interesting example of hardness. In graphite, (the "lead" found in pencils) the carbon is very soft, while the carbon in a diamond is roughly seven times as hard.

Figure 3.5.1: Pencil (left) and Diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

Melting and boiling points are somewhat unique identifiers, especially of compounds. In addition to giving some idea as to the identity of the compound, important information can be obtained about the purity of the material.

Chemical Properties

Chemical properties of matter describe its potential to undergo some chemical change or reaction by virtue of its composition. The elements, electrons, and bonds that are present give the matter potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, after studying chemistry for some time, you should be able to look at the formula of a compound and state some chemical property. For example, hydrogen has the potential to ignite and explode given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.

Figure 3.5.2: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure 3.5.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 3.5.2: Contrasting Physical and Chemical Properties

Physical Properties	Chemical Properties
Gallium metal melts at 30 ^{o}C.	Iron metal rusts.
Mercury is a very dense liquid.	A green banana turns yellow when it ripens.
Gold is shiny.	A dry piece of paper burns.

✓ Example $3.5.1$

Which of the following is a chemical property of iron?

a. Iron corrodes in moist air.
b. Density = 7.874 g/cm^3
c. Iron is soft when pure.
d. Iron melts at 1808 K.

Solution

"Iron corrodes in moist air" is the only chemical property of iron from the list.

? Exercise $3.5.1A$

Which of the following is a physical property of matter?

a. corrosiveness
b. pH (acidity)
c. density
d. flammability

Answer

c

? Exercise $3.5.1B$

Which of the following is a chemical property?

a. flammability
b. melting point
c. boiling point
d. density

Answer

a

Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed.

3.6: Changes in Matter - Physical and Chemical Changes

Learning Objectives

- Label a change as chemical or physical.
- List evidence that can indicate a chemical change occurred.

Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are classified as either physical or chemical changes. Chemists learn a lot about the nature of matter by studying the changes that matter can undergo. Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.

Physical Change

Physical changes are changes in which no bonds are broken or formed. This means that the same types of compounds or elements that were there at the beginning of the change are there at the end of the change. Because the ending materials are the same as the beginning materials, the properties (such as color, boiling point, etc.) will also be the same. Physical changes involve moving molecules around, but not changing them. Some types of physical changes include:

- Changes of state (changes from a solid to a liquid or a gas and vice versa).
- Separation of a mixture.
- Physical deformation (cutting, denting, stretching).
- Making solutions (special kinds of mixtures).

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. Melting is an example of a physical change. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. When liquid water is heated, it changes to water vapor. However, even though the physical properties have changed, the molecules are exactly the same as before. We still have each water molecule containing two hydrogen atoms and one oxygen atom covalently bonded. When you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of the pennies or the nickels—you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you started tearing is still paper when you are done. Again, this is an example of a physical change.

Figure 3.6.1: Ice melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; however, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.

Chemical Change

Chemical changes occur when bonds are broken and/or formed between molecules or atoms. This means that one substance with a certain set of properties (such as melting point, color, taste, etc) is turned into a different substance with different properties. Chemical changes are frequently harder to reverse than physical changes.

One good example of a chemical change is burning a candle. The act of burning paper actually results in the formation of new chemicals (carbon dioxide and water) from the burning of the wax. Another example of a chemical change is what occurs when natural gas is burned in your furnace. This time, on the left there is a molecule of methane, CH_4, and two molecules of oxygen, O_2; on the right are two molecules of water, H_2O, and one molecule of carbon dioxide, CO_2. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.

Figure 3.6.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

We can't actually see molecules breaking and forming bonds, although that's what defines chemical changes. We have to make other observations to indicate that a chemical change has happened. Some of the evidence for chemical change will involve the energy changes that occur in chemical changes, but some evidence involves the fact that new substances with different properties are formed in a chemical change.

Observations that help to indicate chemical change include:

- Temperature changes (either the temperature increases or decreases).
- Light given off.
- Unexpected color changes (a substance with a different color is made, rather than just mixing the original colors together).
- Bubbles are formed (but the substance is not boiling—you made a substance that is a gas at the temperature of the beginning materials, instead of a liquid).
- Different smell or taste (do not taste your chemistry experiments, though!).
- A solid forms if two clear liquids are mixed (look for *floaties*—technically called a precipitate).

✓ Example 3.6.1

Label each of the following changes as a physical or chemical change. Give evidence to support your answer.

a. Boiling water.
b. A nail rusting.
c. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
d. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

a. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $H_2O\,(l) \rightarrow H_2O\,(g)$
b. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $Fe + O_2 \rightarrow Fe_2O_3$

c. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
d. Chemical: the formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

? Exercise 3.6.1

Label each of the following changes as a physical or chemical change.

a. A mirror is broken.
b. An iron nail corroded in moist air
c. Copper metal is melted.
d. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer a:

physical change

Answer b:

chemical change

Answer c:

physical change

Answer d:

chemical change

Separating Mixtures Through Physical Changes

Homogeneous mixtures (solutions) can be separated into their component substances by physical processes that rely on differences in some physical property, such as differences in their boiling points. Two of these separation methods are distillation and crystallization. Distillation makes use of differences in volatility, a measure of how easily a substance is converted to a gas at a given temperature. A simple distillation apparatus for separating a mixture of substances, at least one of which is a liquid. The most volatile component boils first and is condensed back to a liquid in the water-cooled condenser, from which it flows into the receiving flask. If a solution of salt and water is distilled, for example, the more volatile component, pure water, collects in the receiving flask, while the salt remains in the distillation flask.

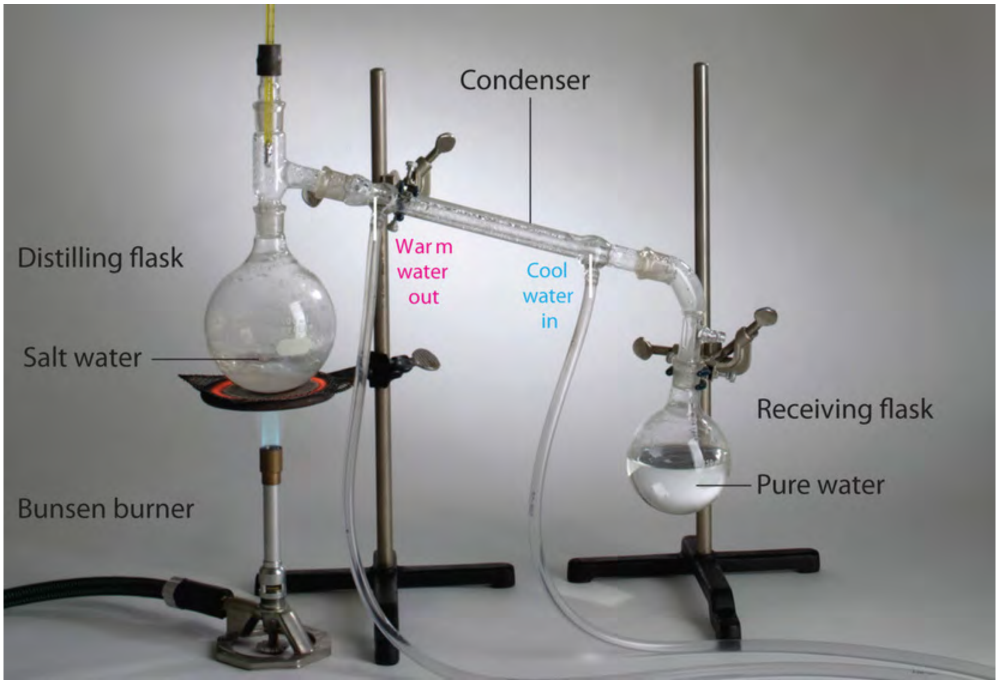

Figure $3.6.3$: *The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.*

Parts of a distillation setup: Bunsen burner, salt water in distilling flask, condenser with cool water in and warm water out, pure water in receiving flask

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as brandy or whiskey. This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States.

Another example for using physical properties to separate mixtures is filtration (Figure $3.6.4$). Filtration is any mechanical, physical or biological operation that separates solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.

Figure 3.6.4: *Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).*

Summary

- Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.
- Physical changes are changes that do not alter the identity of a substance.
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change has occurred include color change, temperature change, light given off, formation of bubbles, formation of a precipitate, etc.

Contributions & Attributions

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- Boundless (www.boundless.com)

3.7: Conservation of Mass - There is No New Matter

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Look at Figure 3.7.1 below. It shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose you had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose you had been able to measure the oxygen used by the fire and the gases produced by the fire. What would you find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

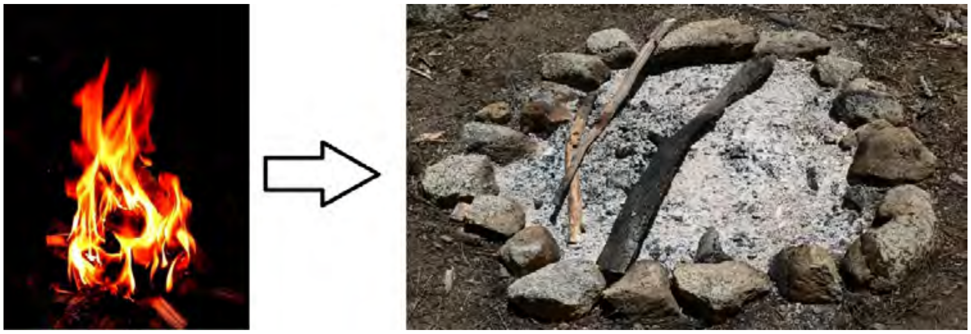

Figure 3.7.1: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). (CC BY-SA 2.5; Einar Helland Berger for fire and Walter Siegmund for ash).

Law of Conservation of Mass

The law of conservation of mass was created in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases equals the original mass of the charcoal and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. A reactant is the chemical reaction of two or more elements to make a new substance, and a product is the substance that is formed as the result of a chemical reaction (Video 3.7.1). Matter and its corresponding mass may not be able to be created or destroyed, but can change forms to other substances like liquids, gases, and solids.

Video 3.7.1: This is a nice little demonstration showing the Conservation of Mass in action.

If you witness a 300 kg tree burn to the ground, there are only ashes left after the burn, and all of them together weigh 10 kg. It may make you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only thing left that you can see is the 10 kg of ash. If you know the law of conservation of mass, then you know that the other 290 kg has to go somewhere, because it has to equal the mass of the tree before it burnt down.

✓ Example 3.7.1

If heating 10.0 grams of calcium carbonate ($CaCO_3$) produces 4.4 g of carbon dioxide (CO_2) and 5.6 g of calcium oxide (CaO), show that these observations are in agreement with the law of conservation of mass.

Solution

$$\begin{aligned}\text{Mass of the reactants} &= \text{Mass of the products}\\ 10.0\text{ g of }CaCO_3 &= 4.4\text{ g of }CO_2 + 5.6\text{ g of }CaO\\ 10.0\text{ g of reactant} &= 10.0\text{ g of products}\end{aligned}$$

Because the mass of the reactant is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

? Exercise 3.7.1

Potassium hydroxide (KOH) readily reacts with carbon dioxide (CO_2) to produce potassium carbonate (K_2CO_3) and water (H_2O). How many grams of potassium carbonate are produced if 224.4 g of KOH reacts with 88.0 g of CO_2? The reaction also produces 36.0 g of water.

Answer

276.4 g of potassium carbonate

The Law is also applicable to both chemical and physical changes. For example, if you have an ice cube that melts into a liquid and you heat that liquid up, it becomes a gas. It will appear to have disappeared, but is still there.

Summary

- Burning and other changes in matter do not destroy matter.
- The mass of matter is always the same before and after the changes occur.
- The law of conservation of mass states that matter cannot be created or destroyed.

Contributions & Attributions

- Binod Shrestha (University of Lorraine)

3.8: Energy

Learning Objectives

- Define heat and work.
- Distinguish between kinetic energy and potential energy.
- State the law of conservation of matter and energy.

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we are awake or asleep.

Ability to Do Work or Produce Heat

When we speak of *using* energy, we are really referring to *transferring* energy from one place to another. When you use energy to throw a ball, you transfer energy from your body to the ball, and this causes the ball to fly through the air. When you use energy to warm your house, you transfer energy from the furnace to the air in your home, and this causes the temperature in your house to rise. Although energy is used in many kinds of different situations, all of these uses rely on energy being transferred in one of two ways. Energy can be transferred as *heat* or as *work*.

When scientists speak of ***heat***, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature, as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot—the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure 3.8.1).

Figure 3.8.1: Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the food.

Heat is only one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is *force (any push or pull) applied over a distance*. When you push an object and cause it to move, you do work, and you transfer some of *your* energy to the object. At this point, it's important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but it is not true in science. By definition, scientific work requires that force be applied *over a distance*. It does not matter how hard you push or how hard you pull. If you have not moved the object, you haven't done any work.

So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work. But the question still remains—*what IS energy*?

Kinetic Energy

Machines use energy, our bodies use energy, energy comes from the sun, energy comes from volcanoes, energy causes forest fires, and energy helps us to grow food. With all of these seemingly different types of energy, it's hard to believe that there are really only two different *forms* of energy: kinetic energy and potential energy. **Kinetic energy** is energy associated with motion. When an object is moving, it has kinetic energy. When the object stops moving, it has no kinetic energy. While all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you are from California, you might have driven through the Tehachapi Pass near Mojave or the Montezuma Hills in Solano County and seen the windmills lining the slopes of the mountains (Figure $3.8.2$). These are two of the larger wind farms in North America. As wind rushes along the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.

Figure $3.8.2$: A wind farm in Solano County harnesses the kinetic energy of the wind. (CC BY-SA 3.0 Unported; BDS2006 at Wikipedia)

Potential Energy

Potential energy is *stored* energy. It is energy that remains available until we choose to use it. Think of a battery in a flashlight. If left on, the flashlight battery will run out of energy within a couple of hours, and the flashlight will die. If, however, you only use the flashlight when you need it, and turn it off when you don't, the battery will last for days or even months. The battery contains a certain amount of energy, and it will power the flashlight for a certain amount of time, but because the battery stores *potential* energy, you can choose to use the energy all at once, or you can save it and only use a small amount at a time.

Any stored energy is potential energy. There are a lot of different ways in which energy can be stored, and this can make potential energy very difficult to recognize. In general, an object has potential energy because of its *position relative to another object*. For example, when a rock is held above the earth, it has potential energy because of its position relative to the ground. This is *potential energy* because the energy is *stored* for as long as the rock is held in the air. Once the rock is dropped, though, the stored energy is released as kinetic energy as the rock falls.

Chemical Energy

There are other common examples of potential energy. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When two magnets are held next to one another, they store potential energy too. For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

As you learned earlier, the world, and all of the chemicals in it are made up of atoms and molecules. These store potential energy that is dependent on their positions relative to one another. Of course, you can't see atoms and molecules. Nevertheless, scientists do know a lot about the ways in which atoms and molecules interact, and this allows them to figure out how much potential energy is stored in a specific quantity (like a cup or a gallon) of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

> Since different chemicals have different amounts of potential energy, scientists will sometimes say that potential energy depends not only on ***position***, but also on ***composition***. Composition affects potential energy because it determines which

> molecules and atoms end up next to one another. For example, the total potential energy in a cup of pure water is different than the total potential energy in a cup of apple juice, because the cup of water and the cup of apple juice are *composed* of different amounts of different chemicals.

At this point, you may wonder just how useful chemical potential energy is. If you want to release the potential energy stored in an object held above the ground, you just drop it. But how do you get potential energy out of chemicals? It's actually not difficult. Use the fact that different chemicals have *different amounts of potential energy*. If you start with chemicals that have a lot of potential energy and allow them to react and form chemicals with less potential energy, all the extra energy that was in the chemicals at the beginning, but not at the end, is released.

Units of Energy

Energy is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by $1^{\mathrm{o}}\mathrm{C}$. For example, raising the temperature of $100\ \mathrm{g}$ of water from $20^{\mathrm{o}}\mathrm{C}$ to $22^{\mathrm{o}}\mathrm{C}$ would require $100 \times 2 = 200$ cal.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains $85\ \mathrm{kcal}$ or $85,000\ \mathrm{cal}$ In order to make the distinction, the dietary calorie is written with a capital C.

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories}$$

To say that the snack "contains" 85 Calories means that $85\ \mathrm{kcal}$ of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1\ \mathrm{J} = 0.2390\ \mathrm{cal}\ \text{or}\ 1\ \mathrm{cal} = 4.184\ \mathrm{J}$$

We can calculate the amount of heat released in kilojoules when a 400 Calorie hamburger is digested.

$$400\ \mathrm{Cal} = 400\ \mathrm{kcal} \times \frac{4.184\ \mathrm{kJ}}{1\ \mathrm{kcal}} = 1.67 \times 10^{3}\ \mathrm{kJ}$$

Summary

- Any time we use energy, we transfer energy from one object to another.
- Energy can be transferred in one of two ways: as heat, or as work.
- Heat is the term given to energy that is transferred from a hot object to a cooler object due to the difference in their temperatures.
- Work is the term given to energy that is transferred as a result of a force applied over a distance.
- Energy comes in two fundamentally different forms: kinetic energy and potential energy.
- Kinetic energy is the energy of motion.
- Potential energy is stored energy that depends on the position of an object relative to another object.
- Chemical potential energy is a special type of potential energy that depends on the positions of different atoms and molecules relative to one another.
- Chemical potential energy can also be thought of according to its dependence on chemical composition.
- Energy can be converted from one form to another. The total amount of mass and energy in the universe is conserved.

Contributions & Attributions

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- Wikibooks

3.9: Energy and Chemical and Physical Change

Learning Objectives

- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

So far, we have talked about how energy exists as either *kinetic energy* or *potential energy* and how energy can be transferred as either *heat* or *work*. While it's important to understand the difference between kinetic energy and potential energy and the difference between heat and work, the truth is, energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Likewise, energy that is transferred as work might later end up transferred as heat, while energy that is transferred as heat might later end up being used to do work.

Even though energy can change form, it must still follow one fundamental law: *Energy cannot be created or destroyed, it can only be changed from one form to another*. This law is known as the **Law of Conservation of Energy**. In a lot of ways, energy is like money. You can exchange quarters for dollar bills and dollar bills for quarters, but no matter how often you convert between the two, you will not end up with any more or any less money than you started with. Similarly, you can transfer (or spend) money using cash, or transfer money using a credit card; but you still spend the same amount of money, and the store still makes the same amount of money.

A campfire is an example of basic thermochemistry. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, roast marshmallows, or just keep warm when it's cold outside.

An image of a campfire with colored flames, made by the burning of a garden hose in a copper pipe. (CC SA-BY 3.0; Jared)

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe: *the system* and *the surroundings*. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** are everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings are the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

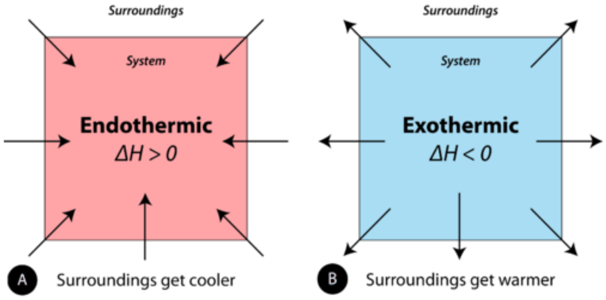

Figure 3.9.1: (A) Endothermic reaction. (B) Exothermic reaction.

Endothermic reaction: surroundings get cooler and delta H is greater than 0, Exothermic reaction: surroundings get warmer and delta H is less than 0

During phase changes, energy changes are usually involved. For example, when solid dry ice vaporizes (physical change), carbon dioxide molecules absorb energy. When liquid water becomes ice, energy is released. Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have **potential energy** or "stored energy". Because we are changing the bonding, this means we are also changing how much of this "stored energy" there is in a reaction.

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the stored/hidden energy of the reactants and products as well as the activation energy. If, on an energy diagram, the products have more stored energy than the reactants started with, the reaction is endothermic. You had to give the reaction energy. If, on the energy diagram, the products have less stored energy than the reactants started with, the reaction is exothermic.

✓ Example 3.9.1

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. gasoline burning
c. ice forming on a pond

Solution

a. Endothermic—you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b. Exothermic—when you burn something, it feels hot to you because it is giving off heat into the surroundings.
c. Exothermic—think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

? Exercise 3.9.1

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer (a)

exothermic

Answer (b)

endothermic

Summary

Phase changes involve changes in energy. All chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy. Reactions that absorb energy are endothermic. Reactions that release energy are exothermic.

3.10: Temperature - Random Motion of Molecules and Atoms

Learning Objectives

- Identify the different between temperature and heat.
- Recognize the different scales used to measure temperature

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.

Figure 3.10.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.

Figure 3.10.2: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as 32^oF and the boiling point as 212^oF. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at 0^oC and 100^oC respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".

The Kelvin Scale

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of $0\ \mathrm{K}$, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is $273.15\ \mathrm{K}$, while the boiling point is $373.15\ \mathrm{K}$ Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales where temperatures are referred to as "degrees F" or "degrees C", we simply designate temperatures in the Kelvin scale as kelvins.

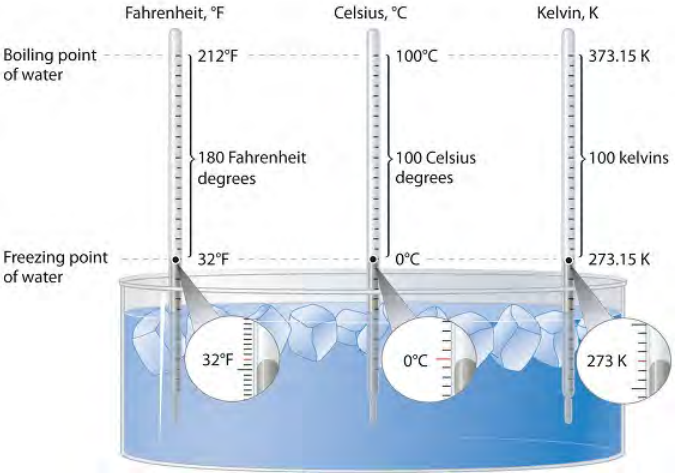

Figure 3.10.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Converting Between Scales

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in }^\circ\mathrm{C}) + 273.15 = T\,(\text{in K}) \tag{3.10.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in }^\circ\mathrm{C}) \tag{3.10.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales is as follows:

$$^\circ C = \frac{(^\circ F - 32)}{1.8} \tag{3.10.3}$$

$$^\circ F = 1.8 \times (^\circ C) + 32 \tag{3.10.4}$$

✓ Example 3.10.1: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 3.10.3:

$$^\circ C = \frac{(103.5\,^\circ F - 32)}{1.8} \tag{3.10.1}$$

$$= 39.7\,^\circ C \tag{3.10.2}$$

Converting from Celsius to Kelvin requires the use of Equation 3.10.1:

$$K = 39.7\,^{\circ}C + 273.15 \tag{3.10.3}$$
$$= 312.9\,K \tag{3.10.4}$$

? Exercise 3.10.1

Convert each temperature to °C and °F.

a. the temperature of the surface of the sun (5800 K)
b. the boiling point of gold (3080 K)
c. the boiling point of liquid nitrogen (77.36 K)

Answer (a)

5527 K, 9980 °F

Answer (b)

2807 K, 5084 °F

Answer (c)

-195.79 K, -320.42 °F

Summary

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).

Contributions & Attributions

3.11: Temperature Changes - Heat Capacity

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. The heat capacity of an object depends on both its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the wading pool.

Heat Capacity and Specific Heat

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot. We would say that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by 1^oC). Water is very resistant to changes in temperature, while metals in general are not. The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1^oC. The symbol for specific heat is c_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree $(\text{J/g}^\text{o}\text{C})$ or calories per gram per degree $(\text{cal/g}^\text{o}\text{C})$ (Table $3.11.1$). This text will use $\text{J/g}^\text{o}\text{C}$ for specific heat.

$$\text{specific heat} = \frac{\text{heat}}{\text{mass} \times \text{cal/g}^\text{o}\text{C}}$$

Notice that water has a very high specific heat compared to most other substances.

Table 3.11.1: Specific Heat Capacities

Substance	Specific Heat Capacity at 25$^\text{o}$C in J/g $^\text{o}$C	Substance	Specific Heat Capacity at 25$^\text{o}$C in J/g $^\text{o}$C
H_2 gas	14.267	steam @ 100$^\text{o}$C	2.010
He gas	5.300	vegetable oil	2.000
$H_2O(l)$	4.184	sodium	1.23
lithium	3.56	air	1.020
ethyl alcohol	2.460	magnesium	1.020
ethylene glycol	2.200	aluminum	0.900
ice @ 0$^\text{o}$C	2.010	concrete	0.880
steam @ 100$^\text{o}$C	2.010	glass	0.840

Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see table above). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

Figure $3.11.1$: This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by 1°C).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C.

3.12: Energy and Heat Capacity Calculations

Learning Objectives

- To relate heat transfer to temperature change.

Heat is a familiar manifestation of transferring energy. When we touch a hot object, energy flows from the hot object into our fingers, and we perceive that incoming energy as the object being “hot.” Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as “cold.” In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Every substance has a characteristic specific heat, which is reported in units of cal/g•°C or cal/g•K, depending on the units used to express Δ*T*. The specific heat of a substance is the amount of energy that must be transferred to or from 1 g of that substance to change its temperature by 1°. Table $3.12.1$ lists the specific heats for various materials.

Table 3.12.1*: Specific Heats of Some Common Substances*

Substance	**Specific Heat** $(\text{J/g}^\text{o}\text{C})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

The *direction* of heat flow is not shown in heat = *mc*Δ*T*. If energy goes into an object, the total energy of the object increases, and the values of heat Δ*T* are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and Δ*T* are negative.

✓ Example $3.12.1$

A $15.0\,\text{g}$ piece of cadmium metal absorbs $134\,\text{J}$ of heat while rising from 24.0^oC to 62.7^oC. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Heat} = q = 134 \text{ J}$
- $\text{Mass} = m = 15.0 \text{ g}$
- $\Delta T = 62.7^\text{o}\text{C} - 24.0^\text{o}\text{C} = 38.7^\text{o}\text{C}$

Unknown

- c_p of cadmium $= ? \text{ J/g}^\text{o}\text{C}$

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^\text{o}\text{C}} = 0.231 \text{ J/g}^\text{o}\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known (Table $3.12.1$), they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g of water at 23.52^oC was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813 \text{ J}}{4.18 \text{ J/g}^\text{o}\text{C} \times 60.0 \text{ g}} = 3.24^\text{o}\text{C}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^\text{o}\text{C} - 3.24^\text{o}\text{C} = 20.28^\text{o}\text{C}$$

✓ Example $3.12.2$

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Solution

We can use heat = $mc\Delta T$ to determine the amount of heat, but first we need to determine ΔT. Because the final temperature of the iron is 73.3°C and the initial temperature is 25.0°C, ΔT is as follows:

$$\Delta T = T_\text{final} - T_\text{initial} = 73.3°\text{C} - 25.0°\text{C} = 48.3°\text{C}$$

The mass is given as 150.0 g, and Table 7.3 gives the specific heat of iron as 0.108 cal/g•°C. Substitute the known values into heat = $mc\Delta T$ and solve for amount of heat:

$$\text{heat} = (150.0 \text{ g}) \left(0.108 \frac{\text{cal}}{\text{g} \cdot {}^\circ\text{C}}\right) (48.3^\circ\text{C}) = 782 \text{ cal}$$

Note how the gram and °C units cancel algebraically, leaving only the calorie unit, which is a unit of heat. Because the temperature of the iron increases, energy (as heat) must be flowing *into* the metal.

? Exercise 3.12.1

What quantity of heat is transferred when a 295.5 g block of aluminum metal is cooled from 128.0°C to 22.5°C? What is the direction of heat flow?

Answer

Heat leaves the aluminum block.

✓ Example 3.12.2

A 10.3 g sample of a reddish-brown metal gave off 71.7 cal of heat as its temperature decreased from 97.5°C to 22.0°C. What is the specific heat of the metal? Can you identify the metal from the data in Table 3.12.1?

Solution

The question gives us the heat, the final and initial temperatures, and the mass of the sample. The value of ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 22.0°\text{C} - 97.5°\text{C} = -75.5°\text{C}$$

If the sample gives off 71.7 cal, it loses energy (as heat), so the value of heat is written as a negative number, −71.7 cal. Substitute the known values into heat = $mc\Delta T$ and solve for c:

$$-71.7 \text{ cal} = (10.3 \text{ g})(c)(-75.5°\text{C})$$

$$c = \frac{-71.7 \text{ cal}}{(10.3 \text{ g})(-75.5°\text{C})}$$

$$c = 0.0923 \text{ cal/g}\bullet°\text{C}$$

This value for specific heat is very close to that given for copper in Table 7.3.

? Exercise 3.12.2

A 10.7 g crystal of sodium chloride (NaCl) has an initial temperature of 37.0°C. What is the final temperature of the crystal if 147 cal of heat were supplied to it?

Answer

Summary

Specific heat calculations are illustrated.

3.13: Phase Changes

Learning Objectives

- Determine the heat associated with a phase change.

Matter can exist in one of several different states, including a gas, liquid, or solid state. The amount of energy in molecules of matter determines the **state of matter**.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

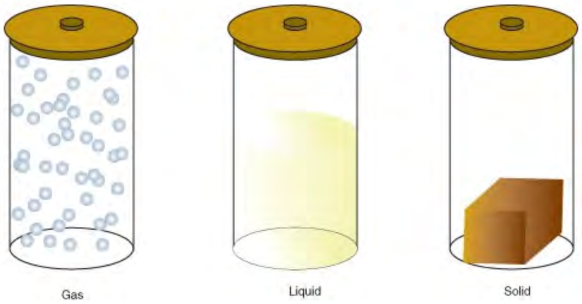

Figure 3.13.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

The following are the changes of state:

Changes of State

Solid → Liquid	Melting or fusion
Liquid → Gas	Vaporization
Liquid → Solid	Freezing
Gas → Liquid	Condensation
Solid → Gas	Sublimation

- If heat is added to a substance, such as in melting, vaporization, and sublimation, the process is **endothermic**. In this instance, heat is increasing the speed of the molecules causing them move faster (examples: solid to liquid; liquid to gas; solid to gas).
- If heat is removed from a substance, such as in freezing and condensation, then the process is **exothermic**. In this instance, heat is decreasing the speed of the molecules causing them move slower (examples: liquid to solid; gas to liquid). These changes **release heat** to the surroundings.
- The amount of heat needed to change a sample from solid to liquid would be the same to reverse from liquid to solid. The only difference is the direction of heat transfer.

Example 3.13.1

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. ice forming on a pond

Solution

a. endothermic - you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b. exothermic - think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

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Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer

a. exothermic

b. endothermic

A phase change is a physical process in which a substance goes from one phase to another. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance. The melting point is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid). The boiling point is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid). The nature of the phase change depends on the direction of the heat transfer. Heat going *into* a substance changes it from a solid to a liquid or a liquid to a gas. Removing heat *from* a substance changes a gas to a liquid or a liquid to a solid.

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O. If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O. A similar process can occur at 100°C: adding heat increases the amount of gaseous H_2O, while removing heat increases the amount of liquid H_2O (Figure 3.13.1).

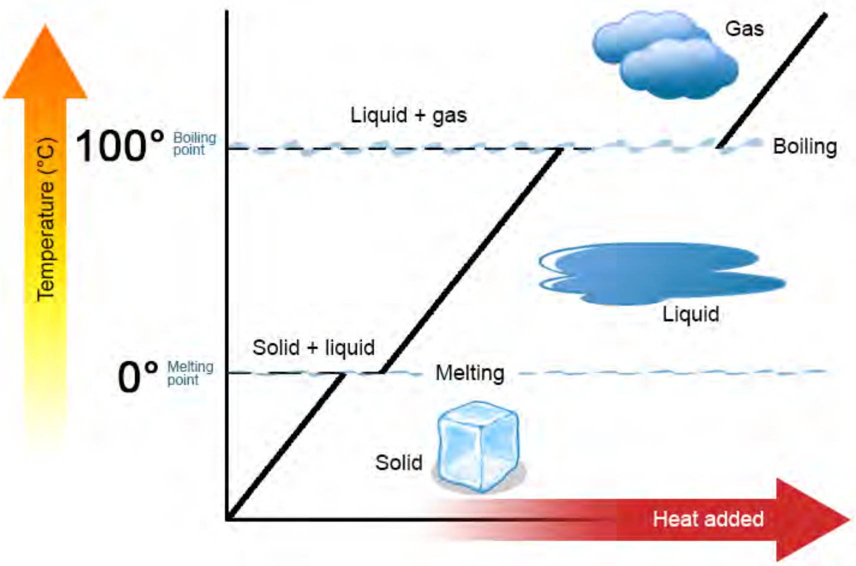


Figure 3.13.2: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).

> Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well.

Second, as shown in Figure $3.13.1$, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means “constant temperature”). Again, consider H_2O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H_2O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

For each phase change of a substance, there is a characteristic quantity of heat needed to perform the phase change per gram (or per mole) of material. The heat of fusion (ΔH_{fus}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the melting point. The heat of vaporization (ΔH_{vap}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the boiling point. If you know the total number of grams or moles of material, you can use the ΔH_{fus} or the ΔH_{vap} to determine the total heat being transferred for melting or solidification using these expressions:

$$\text{heat} = n \times \Delta H_{fus} \tag{3.13.1}$$

where n is the number of moles and ΔH_{fus} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{fus} \tag{3.13.2}$$

where m is the mass in grams and ΔH_{fus} is expressed in energy/gram.

For the boiling or condensation, use these expressions:

$$\text{heat} = n \times \Delta H_{vap} \tag{3.13.3}$$

where n is the number of moles) and ΔH_{vap} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{vap} \tag{3.13.4}$$

where m is the mass in grams and ΔH_{vap} is expressed in energy/gram.

Remember that a phase change depends on the direction of the heat transfer. If heat transfers in, solids become liquids, and liquids become solids at the melting and boiling points, respectively. If heat transfers out, liquids solidify, and gases condense into liquids. At these points, there are no changes in temperature as reflected in the above equations.

✓ Example 3.13.2

How much heat is necessary to melt 55.8 g of ice (solid H_2O) at 0°C? The heat of fusion of H_2O is 79.9 cal/g.

Solution

We can use the relationship between heat and the heat of fusion (Equation $3.13.1$) to determine how many cal of heat are needed to melt this ice:

$$\text{heat } = \mathrm{m} \times \Delta \mathrm{H}_{\text{fus}}$$

$$\text{heat } = (55.8 \ \cancel{\mathrm{g}}) \left(\frac{79.9 \text{ cal}}{\cancel{\mathrm{g}}}\right) = 4,460 \text{ cal}$$

? Exercise 3.13.2

How much heat is necessary to vaporize 685 g of H_2O at 100°C? The heat of vaporization of H_2O is 540 cal/g.

Answer

$$\text{heat } = \mathrm{m} \times \Delta \mathrm{H}_{\text{vap}}$$

$$\text{heat } = (685 \ \cancel{\mathrm{g}}) \left(\frac{540 \text{ cal}}{\cancel{\mathrm{g}}}\right) = 370,000 \text{ cal}$$

Table 3.13.1 lists the heats of fusion and vaporization for some common substances. Note the units on these quantities; when you use these values in problem solving, make sure that the other variables in your calculation are expressed in units consistent with the units in the specific heats or the heats of fusion and vaporization.

Table 3.13.1: Heats of Fusion and Vaporization for Selected Substances

Substance	ΔH_{fus} (cal/g)	ΔH_{vap} (cal/g)
aluminum (Al)	94.0	2,602
gold (Au)	15.3	409
iron (Fe)	63.2	1,504
water (H_2O)	79.9	540
sodium chloride (NaCl)	123.5	691
ethanol (C_2H_5OH)	45.2	200.3
benzene (C_6H_6)	30.4	94.1

Sublimation

There is also a phase change where a solid goes directly to a gas:

$$\text{solid} \rightarrow \text{gas} \tag{3.13.5}$$

This phase change is called *sublimation*. Each substance has a characteristic heat of sublimation associated with this process. For example, the heat of sublimation (ΔH_{sub}) of H_2O is 620 cal/g.

We encounter sublimation in several ways. You may already be familiar with dry ice, which is simply solid carbon dioxide (CO_2). At −78.5°C (−109°F), solid carbon dioxide sublimes, changing directly from the solid phase to the gas phase:

$$CO_2(s) \xrightarrow{-78.5^\circ C} CO_2(g) \tag{3.13.6}$$

Solid carbon dioxide is called dry ice because it does not pass through the liquid phase. Instead, it does directly to the gas phase. (Carbon dioxide *can* exist as liquid but only under high pressure.) Dry ice has many practical uses, including the long-term preservation of medical samples.

Even at temperatures below 0°C, solid H_2O will slowly sublime. For example, a thin layer of snow or frost on the ground may slowly disappear as the solid H_2O sublimes, even though the outside temperature may be below the freezing point of water. Similarly, ice cubes in a freezer may get smaller over time. Although frozen, the solid water slowly sublimes, redepositing on the colder cooling elements of the freezer, which necessitates periodic defrosting (frost-free freezers minimize this redeposition). Lowering the temperature in a freezer will reduce the need to defrost as often.

Under similar circumstances, water will also sublime from frozen foods (e.g., meats or vegetables), giving them an unattractive, mottled appearance called freezer burn. It is not really a “burn,” and the food has not necessarily gone bad, although it looks unappetizing. Freezer burn can be minimized by lowering a freezer’s temperature and by wrapping foods tightly so water does not have any space to sublime into.

Key Takeaway

- There is an energy change associated with any phase change.

3.14: What Are Nutrients?

Learning Objective

- Define the word “nutrient” and identify the six classes of nutrients essential for health.
- List the three main energy (Calorie) yielding nutrients and how many Calories each of these nutrients provide.

What is in Food?

Your "diet" is defined as the foods you choose to eat. The diet of most people contains a variety of foods: nuts, seeds, leafy green vegetables and more! The foods we eat contain nutrients. Nutrients are substances required by the body to perform its basic functions. Nutrients must be obtained from diet, since the human body does not synthesize them, or does not synthesize them in large enough amounts for human health. Nutrients are used for many body functions such as: growing, moving your muscles, repairing tissues and much more! There are six classes of essential nutrients required for the body to function and maintain overall health. These six classes of essential nutrients are: carbohydrates, lipids (fats), proteins, water, vitamins, and minerals. Foods also contain non-nutrients. Some non-nutrients appear to be very important for human health, like fiber and antioxidants, some non-nutrients may be harmful to human health such as: preservatives, colorings, flavorings and pesticide residues.

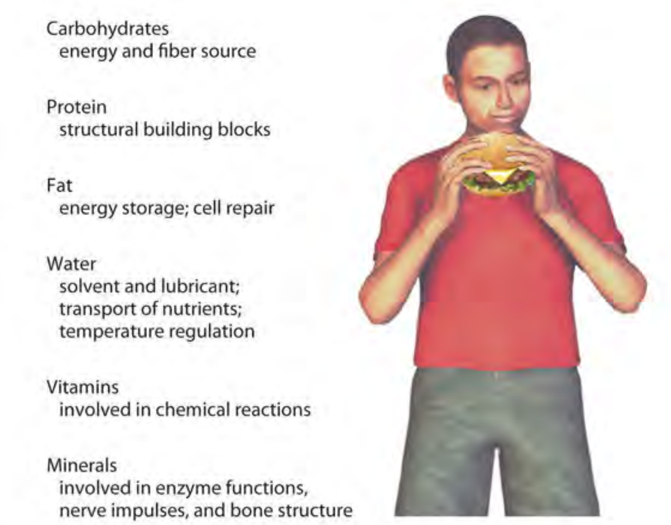


Figure $3.14.1$: The Six Classes of Nutrients. Source: http://www.chemistry.wustl.edu/~edud.../vitamins.html.

Macronutrients

Nutrients that are needed in large amounts are called macronutrients. There are three major classes of macronutrients: carbohydrates, lipids, and proteins. All three of these nutrients are needed in relatively large amounts AND they contain Calories (note the capital C which indicates kilocalories) which can be "burned" in your body to create energy for your body cells. The energy from these macronutrients comes from their chemical bonds. This chemical energy is converted into cellular energy that is then utilized to perform work, allowing our bodies to conduct their basic functions. A unit of measurement of food energy is the Calorie. The next time you are eating a packaged food, look on the "Nutrition Facts" panel to find out how many Calories you getting when you eat one serving of that food. Water is also a macronutrient in the sense that you require a large amount of it, but unlike the other macronutrients it does not yield calories. One other possible component of the diet that can provide Calories is alcohol though it is generally NOT considered to be a nutrient. Alcohol (in the form of ethanol) provides about 7 Calories per gram.

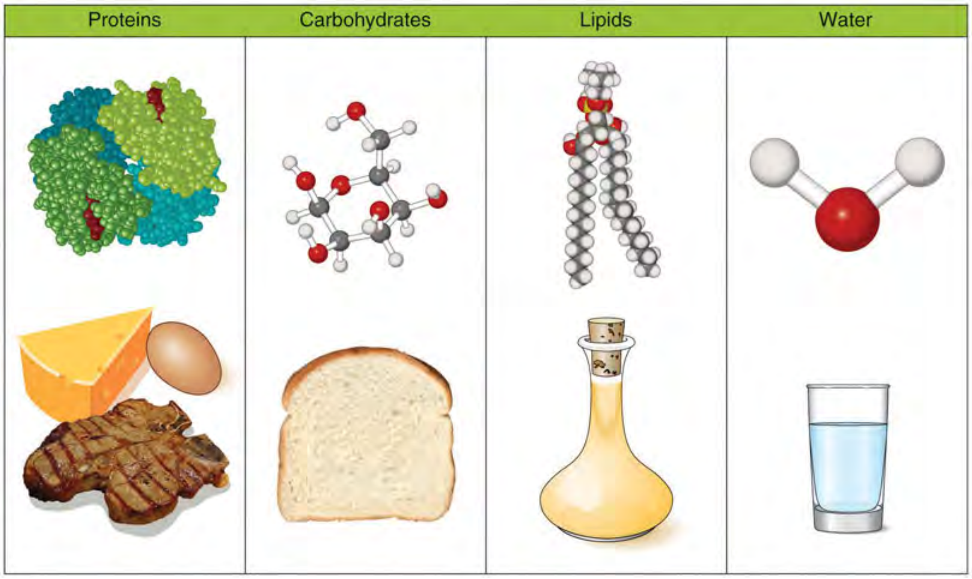


Figure 3.14.2 The Macronutrients: Carbohydrates, Lipids, Protein, and Water

Carbohydrates

When you eat a food that contains carbohydrate, like bread for example, you will be receiving approximately 4 Calories for every gram of carbohydrate you eat. Carbohydrates are molecules composed of carbon, hydrogen, and oxygen. The major food sources of carbohydrates are grains, milk, fruits, and starchy vegetables like potatoes. Non-starchy vegetables also contain carbohydrates, but in lesser quantities. Carbohydrates are broadly classified into two forms based on their chemical structure: fast-releasing carbohydrates, often called "simple sugars", and slow-releasing carbohydrates, often called "complex carbohydrates".

Fast-releasing carbohydrates consist of one or two basic sugar units. They are sometimes called "simple sugars" because their chemical structure is fairly simple with only one or two sugar units. If the carbohydrate has one sugar unit we call it a "monosaccharide". Mono means one and saccharide means sugar. Examples of monosaccharides you have heard of are: Glucose (the sugar that is in your blood); Fructose (a sugar commonly found in fruit); and Galactose (a sugar that is found as part of milk sugar). If the carbohydrate has two sugar units we call it a "disaccharide". Di means two and saccharide means sugar. Examples of disaccharides you may be familiar with are: sucrose (table sugar); lactose (milk sugar) and maltose (malt sugar). Sucrose is made of a Glucose linked together with a Fructose. Lactose is made of Glucose linked together with Galactose. Maltose is made of two Glucose molecules bonded together. In order to digest a disaccharide, your body has to break the two sugars apart. We will talk more about this when we discuss lactose intolerance!

Slow-releasing carbohydrates are long chains of simple sugars (polysaccharides) that can be branched or unbranched. Some polysaccharides that you have probably heard of are starch and fiber. During digestion, the body does its best to break down all slow-releasing carbohydrates like starch to simple sugars, mostly glucose. Glucose is then transported to all our cells where it is stored, used to make energy, or used to build macromolecules. Fiber is also a slow-releasing carbohydrate, but it cannot be broken down in the human body and passes through the digestive tract undigested unless the bacteria that live in the large intestine break it down for us.

One gram of carbohydrates yields four Calories of energy for the cells in the body to perform work. In addition to providing energy and serving as building blocks for bigger macromolecules, carbohydrates are essential for proper functioning of the nervous system, heart, and kidneys. As mentioned, glucose can be stored in the body for future use. In humans, the storage molecule of carbohydrates is called glycogen and in plants it is known as starches. Glycogen and starches are slow-releasing carbohydrates.

Lipids

Lipids are also a family of molecules composed of carbon, hydrogen, and oxygen, but unlike carbohydrates, they are insoluble in water. Lipids are found predominately in butter, oils, meats, dairy products, nuts, and seeds, and in many processed foods. The three main types of lipids are triglycerides (triacylglycerols), phospholipids, and sterols. The main job of lipids is to store energy. Lipids provide more energy per gram than carbohydrates (nine Calories per gram of lipids versus four Calories per gram of carbohydrates). In addition to energy storage, lipids serve as cell membranes, surround and protect organs, aid in temperature regulation, and regulate many other functions in the body.

Proteins

Proteins are macromolecules composed of chains of subunits called amino acids. Amino acids are simple subunits composed of carbon, oxygen, hydrogen, and nitrogen. The food sources of proteins are meats, dairy products, seafood, and a variety of different plant-based foods, most notably soy. The word protein comes from a Greek word meaning "of primary importance," which is an apt description of these macronutrients; they are also known colloquially as the "workhorses" of life. Proteins provide four Calories of energy per gram; however providing energy is not protein's most important function. Proteins provide structure to bones, muscles and skin, and play a role in conducting most of the chemical reactions that take place in the body. Scientists estimate that greater than one-hundred thousand different proteins exist within the human body.

Water

There is one other nutrient that we must have in large quantities: water. Water does not contain carbon, but is composed of two hydrogens and one oxygen per molecule of water. Water does not provide any Calories. More than 60 percent of your total body weight is water. Without it, nothing could be transported in or out of the body, chemical reactions would not occur, organs would not be cushioned, and body temperature would fluctuate widely. On average, an adult consumes just over two liters of water per day from food and drink. According to the "rule of threes," a generalization supported by survival experts, a person can survive three minutes without oxygen, three days without water, and three weeks without food. Since water is so critical for life's basic processes, the amount of water input and output is supremely important, a topic we will explore in detail in Chapter 7.

Alcohol (not a nutrient)

Alcoholic drinks are a source of Calories even though they are generally not considered nutrients. Alcohol itself provides approximately 7 Calories for every gram consumed. In addition to alcohol, many alcoholic drinks contain carbohydrate as well.

Micronutrients

Micronutrients are nutrients required by the body in lesser amounts, but are still essential for carrying out bodily functions. Micronutrients include all the essential minerals and vitamins. There are sixteen essential minerals and thirteen vitamins (Tables $3.14.1$ and $3.14.2$ for a complete list and their major functions). In contrast to carbohydrates, lipids, and proteins, micronutrients do not contain Calories. This is often confusing because most people have heard how tired a person will feel if they are low in a micronutrient such as Iron. The tiredness can be explained by the fact that, micronutrients assist in the process of making energy by being part of enzymes (i.e., coenzymes). Enzymes catalyze chemical reactions in the body and are involved in many aspects of body functions from producing energy, to digesting nutrients, to building macromolecules. Micronutrients play many roles in the body.

Minerals

Minerals are solid inorganic substances that form crystals and are classified depending on how much of them we need. Trace minerals, such as molybdenum, selenium, zinc, iron, and iodine, are only required in a few milligrams or less and macrominerals, such as calcium, magnesium, potassium, sodium, and phosphorus, are required in hundreds of milligrams. Many minerals are critical for enzyme function, others are used to maintain fluid balance, build bone tissue, synthesize hormones, transmit nerve impulses, contract and relax muscles, and protect against harmful free radicals.

Table $3.14.1$: : Minerals and Their Major Functions

Minerals	Major Functions
Macro	
Sodium	Fluid balance, nerve transmission, muscle contraction
Chloride	Fluid balance, stomach acid production
Potassium	Fluid balance, nerve transmission, muscle contraction
Calcium	Bone and teeth health maintenance, nerve transmission, muscle contraction, blood clotting

Minerals	Major Functions
Phosphorus	Bone and teeth health maintenance, acid-base balance
Magnesium	Protein production, nerve transmission, muscle contraction
Sulfur	Protein production
Trace	
Iron	Carries oxygen, assists in energy production
Zinc	Protein and DNA production, wound healing, growth, immune system function
Iodine	Thyroid hormone production, growth, metabolism
Selenium	Antioxidant
Copper	Coenzyme, iron metabolism
Manganese	Coenzyme
Fluoride	Bone and teeth health maintenance, tooth decay prevention
Chromium	Assists insulin in glucose metabolism
Molybdenum	Coenzyme

Vitamins

The thirteen vitamins are categorized as either water-soluble or fat-soluble. The water-soluble vitamins are vitamin C and all the B vitamins, which include thiamine, riboflavin, niacin, pantothenic acid, pyroxidine, biotin, folate and cobalamin. The fat-soluble vitamins are A, D, E, and K. Vitamins are required to perform many functions in the body such as making red blood cells, synthesizing bone tissue, and playing a role in normal vision, nervous system function, and immune system function.

Table 3.14.1: Vitamins and Their Major Functions

Vitamins	Major Functions
Water-soluble	
B_1 (thiamine)	Coenzyme, energy metabolism assistance
B_2 (riboflavin)	Coenzyme, energy metabolism assistance
B_3 (niacin)	Coenzyme, energy metabolism assistance
B_5 (pantothenic acid)	Coenzyme, energy metabolism assistance
B_6 (pyroxidine)	Coenzyme, amino acid synthesis assistance
Biotin	Coenzyme
Folate	Coenzyme, essential for growth
B_{12} (cobalamin)	Coenzyme, red blood cell synthesis
C	Collagen synthesis, antioxidant
Fat-soluble	
A	Vision, reproduction, immune system function
D	Bone and teeth health maintenance, immune system function
E	Antioxidant, cell membrane protection
K	Bone and teeth health maintenance, blood clotting

Vitamin deficiencies can cause severe health problems. For example, a deficiency in niacin causes a disease called pellagra, which was common in the early twentieth century in some parts of America. The common signs and symptoms of pellagra are known as the “4D’s—diarrhea, dermatitis, dementia, and death.” Until scientists found out that better diets relieved the signs and symptoms of pellagra, many people with the disease ended up in insane asylums awaiting death (Video 3.14.1). Other vitamins were also found to prevent certain disorders and diseases such as scurvy (vitamin C), night blindness (vitamin A), and rickets (vitamin D).

Video 3.14.1: This video provides a brief history of Dr. Joseph Goldberger's discovery that pellagra was a diet-related disease.

Food Energy

Though this is only Chapter 1, you have already seen the words "Calories" and "Energy" used several times. In everyday life you have probably heard people talk about how many Calories they burned on the treadmill or how many Calories are listed on a bag of chips. Calories, are a measure of energy. It takes quite a lot of Calories (energy) to keep us alive. Even if a person is in a coma, they still burn approximately 1000 Calories of energy in order for: their heart to beat, their blood to circulate, their lungs to breathe, etc... We burn even more calories when we exercise. The carbohydrates, fats and proteins we eat and drink provide calories for us (and alcohol as well if we choose to consume it). Sometimes people refer to these nutrients as "energy yielding". As you read above, carbohydrates provide 4 Calories for every gram we consume; proteins provide 4 Calories for every gram we consume; fats provide 9 Calories for every gram we consume and alcohol provides 7 Calories of energy for every gram we consume. Vitamins, minerals and water do not provide any calories, even though they are still essential nutrients.

Food Quality: Nutrient Density compared to Calorie Density

One way to think about the quality of your food is to consider how many nutrients you receive from a serving of that food compared to the number of Calories you receive from one serving of the food. For example, a candy bar gives you quite a few Calories but not very many of the essential nutrients. We would say that the candy bar has Calorie density but low Nutrient density. An apple on the other hand, has quite a few essential nutrients but not very many Calories. We would say that the apple has high Nutrient density but low Calorie density.

Please watch the following Ted Ed video called "What is a Calorie" which will describe in more detail what a Calorie is and why you may want to know how many Calories you are consuming compared to how many you are burning each day.

Video: What is a calorie? - Emma Bryce

One measurement of food quality is the amount of nutrients it contains relative to the amount of energy (Calories) it provides. High-quality foods are nutrient dense, meaning they contain lots of the nutrients relative to the amount of Calories they provide. Nutrient-dense foods are the opposite of “empty-calorie” foods such as carbonated sugary soft drinks, which provide many calories and very little, if any, other nutrients. Food quality is additionally associated with its taste, texture, appearance, microbial content, and how much consumers like it.

Food: A Better Source of Nutrients

It is better to get all your micronutrients from the foods you eat as opposed to from supplements. Supplements contain only what is listed on the label, but foods contain many more macronutrients, micronutrients, and other chemicals, like antioxidants that benefit health. While vitamins, multivitamins, and supplements are a $20 billion industry in this country and more than 50 percent of Americans purchase and use them daily, there is no consistent evidence that they are better than food in promoting health and preventing disease. Dr. Marian Neuhouser, associate of the Fred Hutchinson Cancer Research Center in Seattle, says that “…scientific data are lacking on the long-term health benefits of supplements. To our surprise, we found that multivitamins did not lower the risk of the most common cancers and also had no impact on heart disease. ”Woodward, K. “Multivitamins Each Day Will Not Keep Common Cancers Away; Largest Study of Its Kind Provides Definitive Evidence that Multivitamins Will Not Reduce Risk of Cancer or Heart Disease in Postmenopausal Women.” Fred Hutchinson Cancer Research Center. *Center News* 16 (February 2009). http://www.fhcrc.org/about/pubs/center_news/online/2009/02/multivitamin_study.html

Key Takeaways

- Foods contain nutrients that are essential for our bodies to function.
- Four of the classes of nutrients required for bodily function are needed in large amounts. They are carbohydrates, lipids, proteins, and water, and are referred to as macronutrients.
- Two of the classes of nutrients are needed in lesser amounts, but are still essential for bodily function. They are vitamins and minerals.
- One measurement of food quality is the amount of essential nutrients a food contains relative to the amount of energy it has (nutrient density).

Discussion Starters

1. Make a list of some of your favorite foods and visit the “What’s In the Foods You Eat?” search tool provided by the USDA. What are some of the nutrients found in your favorite foods?

   http://www.ars.usda.gov/Services/docs.htm?docid=17032

2. Have a discussion in class on the “progression of science” and its significance to human health as depicted in the video on pellagra (Video 3.14.1.

3.15: The Broad Role of Nutritional Science

Learning Objective

- Provide an example of how the scientific method works to promote health and prevent disease.

How to Determine the Health Effects of Food and Nutrients

Similar to the method by which a police detective finally charges a criminal with a crime, nutritional scientists discover the health effects of food and its nutrients by first making an observation. Once observations are made, they come up with a hypothesis, test their hypothesis, and then interpret the results. After this, they gather additional evidence from multiple sources and finally come up with a conclusion on whether the food suspect fits the claim. This organized process of inquiry used in forensic science, nutritional science, and every other science is called the scientific method.

Below is an illustration of the scientific method at work—in this case to prove that iodine is a nutrient. Zimmerman, M.B. "Research on Iodine Deficiency and Goiter in the 19th and Early 20th Centuries." *J Nutr* 138, no. 11 (November 2008): 2060–63. Carpenter, K.J. "David Marine and the Problem of Goiter." *J Nutr* 135, no.4 (April 2005): 675–80. In 1811, French chemist Bernard Courtois was isolating saltpeter for producing gunpowder to be used by Napoleon's army. To carry out this isolation he burned some seaweed and in the process observed an intense violet vapor that crystallized when he exposed it to a cold surface. He sent the violet crystals to an expert on gases, Joseph Gay-Lussac, who identified the crystal as a new element. It was named iodine, the Greek word for violet. The following scientific record is some of what took place in order to conclude that iodine is a nutrient.

- **Observation.** Eating seaweed is a cure for goiter, a gross enlargement of the thyroid gland in the neck.
- **Hypothesis.** In 1813, Swiss physician Jean-Francois Coindet hypothesized that the seaweed contained iodine and he could use just iodine instead of seaweed to treat his patients.
- **Experimental test.** Coindet administered iodine tincture orally to his patients with goiter.
- **Interpret results.** Coindet's iodine treatment was successful.
- **Gathering more evidence.** Many other physicians contributed to the research on iodine deficiency and goiter.
- **Hypothesis.** French chemist Chatin proposed that the low iodine content in food and water of certain areas far away from the ocean were the primary cause of goiter and renounced the theory that goiter was the result of poor hygiene.
- **Experimental test.** In the late 1860s the program, "The stamping-out of goiter," started with people in several villages in France being given iodine tablets.
- **Results.** The program was effective and 80 percent of goitrous children were cured.
- **Hypothesis.** In 1918, Swiss doctor Bayard proposed iodizing salt as a good way to treat areas endemic with goiter.
- **Experimental test.** Iodized salt was transported by mules to a small village at the base of the Matterhorn where more than 75 percent of school children were goitrous. It was given to families to use for six months.
- **Results.** The iodized salt was beneficial in treating goiter in this remote population.
- **Experimental test.** Physician David Marine conducted the first experiment of treating goiter with iodized salt in America in Akron, Ohio.
- **Results.** This study conducted on over four-thousand school children found that iodized salt prevented goiter.
- **Conclusions.** Seven other studies similar to Marine's were conducted in Italy and Switzerland that also demonstrated the effectiveness of iodized salt in treating goiter. In 1924, US public health officials initiated the program of iodizing salt and started eliminating the scourge of goiterism. Today more than 70 percent of American households use iodized salt and many other countries have followed the same public health strategy to reduce the health consequences of iodine deficiency.

Evidence-Based Approach to Nutrition

It took more than one hundred years from iodine's discovery as an effective treatment for goiter until public health programs recognized it as such. Although a lengthy process, the scientific method is a productive way to define essential nutrients and determine their ability to promote health and prevent disease. The scientific method is part of the overall evidence-based approach to designing nutritional guidelines. An evidence-based approach to nutrition includes:Briss, P.A., et al. "Developing an Evidence-Based Guide to Community Preventive Services—Methods." *Am J Prev Med* 18, no. 1S (2000): 35–43., Myers, E. "Systems for Evaluating Nutrition Research for Nutrition Care Guidelines: Do They Apply to Population Dietary Guidelines?" *J Am Diet Assoc* 103, no. 12, supplement 2 (December 2003): 34–41.

- Defining the problem or uncertainty (e.g., the rate of colon cancer is higher in people who eat red meat)

- Formulating it as a question (e.g., Does eating red meat contribute to colon cancer?)
- Setting criteria for quality evidence
- Evaluating the body of evidence
- Summarizing the body of evidence and making decisions
- Specifying the strength of the supporting evidence required to make decisions
- Disseminating the findings

The Food and Nutrition Board of the Institute of Medicine, a nonprofit, nongovernmental organization, constructs its nutrient recommendations (i.e., Dietary Reference Intakes, or DRI) using an evidence-based approach to nutrition. The entire procedure for setting the DRI is documented and made available to the public. The same approach is used by the USDA and HHS, which are departments of the US federal government. The USDA and HHS websites are great tools for discovering ways to optimize health; however, it is important to gather nutrition information from multiple resources as there are often differences in opinion among various scientists and public health organizations. While the new *Dietary Guidelines*, published in 2015, have been well-received by some, there are nongovernmental public health organizations that are convinced that some pieces of the guidelines may be influenced by lobbying groups and/or the food industry. For example, the Harvard School of Public Health (HSPH) feels the government falls short by being "too lax on refined grains."The guidelines recommend getting at least *half* of grains from whole grains—according to the HSPH this still leaves too much consumption of refined grains.

For a list of reliable sources that advocate good nutrition to promote health and prevent disease using evidence-based science see Table $3.15.1$. In subsequent sections, we will further discuss distinguishing criteria that will enable you to wade through misleading nutrition information and instead gather your information from reputable, credible websites and organizations.

Table $3.15.1$: Web Resources for Nutrition and Health

Organization	Website
Governmental	
US Department of Agriculture	http://www.usda.gov/wps/portal/usda/usdahome
USDA Center for Nutrition Policy and Promotion	www.cnpp.usda.gov/
US Department of Health and Human Services	http://www.hhs.gov/
Centers for Disease Control and Prevention	http://www.cdc.gov/
Food and Drug Administration	http://www.fda.gov/
Healthy People	http://www.healthypeople.gov/2020/default.aspx
Office of Disease Prevention and Health Promotion	odphp.osophs.dhhs.gov/
Health Canada	http://www.hc-sc.gc.ca/
International	
World Health Organization	http://www.who.int/en/
Food and Agricultural Organization of the United Nations	www.fao.org/
Nongovernmental	
Harvard School of Public Health	www.hsph.harvard.edu/nutritio...rce/index.html
Mayo Clinic	http://www.mayoclinic.com/
Linus Pauling Institute	http://lpi.oregonstate.edu/
American Society for Nutrition	http://www.nutrition.org/
American Medical Association	http://www.ama-assn.org/
American Diabetes Association	http://www.diabetes.org/
The Academy of Nutrition and Dietetics	http://www.eatright.org/
Institute of Medicine: Food and Nutrition	www.iom.edu/Global/Topics/Food-Nutrition.aspx
Dietitians of Canada	http://www.dietitians.ca/

Types of Scientific Studies

There are many types of scientific studies that can be used to provide supporting evidence for a particular hypothesis. The various types of studies include epidemiological studies, interventional clinical trials, and randomized clinical interventional trials.

Epidemiological studies are observational studies that look for health patterns and are often the front-line studies for public health. The CDC defines epidemiological studies as scientific investigations that define frequency, distribution, and patterns of health events in a population. Thus, these studies describe the occurrence and patterns of health events over time. The goal of an epidemiological study is to find factors associated with an increased risk for a health event, though these sometimes remain elusive. An example of an epidemiological study is the Framingham Heart Study, a project of the National Heart, Lung and Blood Institute and Boston University that has been ongoing since 1948. This study first examined the physical health and lifestyles of 5,209 men and women from the city of Framingham, Massachusetts and has now incorporated data from the children and grandchildren of the original participants. One of the seminal findings of this ambitious study was that higher cholesterol levels in the blood are a risk factor for heart disease. The Framingham Heart Study, a project of the National Heart, Lung, and Blood Institute and Boston University. "History of the Framingham Heart Study." Epidemiological studies are a cornerstone for examining and evaluating public health and some of their advantages are that they can lead to the discovery of disease patterns and risk factors for diseases, and they can be used to predict future healthcare needs and provide information for the design of disease prevention strategies for entire populations. Some shortcomings of epidemiological studies are that investigators cannot control environments and lifestyles, a specific group of people studied may not be an accurate depiction of an entire population, and these types of scientific studies cannot directly determine if one variable causes another. Scientists regard epidemiological studies as a starting place for ideas about what types of lifestyle choice MAY contribute to a disease. They can show a correlation between two things happening but not whether one factor CAUSED the other. My favorite example of this is a thought experiment where a make-believe scientist observes (through epidemiological studies) that grey hair often goes along with Alzheimer's Disease. Wouldn't it be silly for that scientist to predict that grey hair CAUSES Alzheimer's Disease? Could that scientist easily test that silly prediction? Suppose that scientist dyed a group of people's hair grey to see if the color grey truly increased their risk of Alzheimer's Disease. The whole thing is silly but I think you get the point. Epidemiological studies may give us ideas but those ideas need clinical trials to figure out what the patterns mean.

Interventional clinical trial studies are scientific investigations in which a variable is changed between groups of people. When well done, this type of study allows one to determine causal relationships. An example of an interventional clinical trial study is the Dietary Approaches to Stop Hypertension (DASH) trial published in the April 1997 issue of *The New England Journal of Medicine*.Appel, L. J., et al. "A Clinical Trial of the Effects of Dietary Patterns on Blood Pressure.," *N Engl J Med* 336 (April 1997): 1117–24. In this study, 459 people were randomly assigned to three different groups; one was put on an average American control diet, a second was put on a diet rich in fruits and vegetables, and the third was put on a combination diet rich in fruits, vegetables, and low-fat dairy products with reduced saturated and total fat intake. The groups remained on the diets for eight weeks. Blood pressures were measured before starting the diets and after eight weeks. Results of the study showed that the group on the combination diet had significantly lower blood pressure at the end of eight weeks than those who consumed the control diet. The authors concluded that the combination diet is an effective nutritional approach to treat high blood pressure. The attributes of high-quality clinical interventional trial studies are:

- those that include a control group, which does not receive the intervention, to which you can compare the people who receive the intervention being tested;
- those in which the subjects are randomized into the group or intervention group, meaning a given subject has an equal chance of ending up in either the control group or the intervention group. This is done to ensure that any possible confounding variables are likely to be evenly distributed between the control and the intervention groups;
- those studies that include a sufficient number of participants.

What are confounding variables? These are factors other than the one being tested that could influence the results of the study. For instance, in the study we just considered, if one group of adults did less physical activity than the other, then it could be the amount of physical activity rather than the diet being tested that caused the differences in blood pressures among the groups.

The limitations of these types of scientific studies are that they are difficult to carry on for long periods of time, are costly, and require that participants remain compliant with the intervention. Furthermore, it is unethical to study certain interventions. (An example of an unethical intervention would be to advise one group of pregnant mothers to drink alcohol to determine the effects of alcohol intake on pregnancy outcome, because we know that alcohol consumption during pregnancy damages the developing fetus.)

Randomized clinical interventional trial studies are powerful tools to provide supporting evidence for a particular relationship and are considered the “gold standard” of scientific studies. A randomized clinical interventional trial is a study in which participants are assigned by chance to separate groups that compare different treatments. Neither the researchers nor the participants can choose which group a participant is assigned. However, from their limitations it is clear that epidemiological studies complement interventional clinical trial studies and both are necessary to construct strong foundations of scientific evidence for health promotion and disease prevention.

Other scientific studies used to provide supporting evidence for a hypothesis include laboratory studies conducted on animals or cells. An advantage of this type of study is that they typically do not cost as much as human studies and they require less time to conduct. Other advantages are that researchers have more control over the environment and the amount of confounding variables can be significantly reduced. Moreover, animal and cell studies provide a way to study relationships at the molecular level and are also helpful in determining the exact mechanism by which a specific nutrient causes a change in health. The disadvantage of these types of studies are that researchers are not working with whole humans and thus the results may not be relevant. Nevertheless, well-conducted animal and cell studies that can be repeated by multiple researchers and obtain the same conclusion are definitely helpful in building the evidence to support a scientific hypothesis.

Evolving Science

Science is always moving forward, albeit sometimes slowly. One study is not enough to make a guideline or a recommendation or cure a disease. Science is a stepwise process that builds on past evidence and finally culminates into a well-accepted conclusion. Unfortunately, not all scientific conclusions are developed in the interest of human health and it is important to know where a scientific study was conducted and who provided the money. Indeed, just as an air quality study paid for by a tobacco company diminishes its value in the minds of readers, so does one on red meat performed at a laboratory funded by a national beef association.

Science can also be contentious even amongst experts that don’t have any conflicting financial interests. To see scientists debating over the nutritional guidelines, watch Video \(\PageIndex{1}\. Contentious science is actually a good thing as it forces researchers to be of high integrity, well-educated, well-trained, and dedicated. It also instigates public health policy makers to seek out multiple sources of evidence in order to support a new policy. Agreement involving many experts across multiple scientific disciplines is necessary for recommending dietary changes to improve health and prevent disease. Although a somewhat slow process, it is better for our health to allow the evidence to accumulate before incorporating some change in our diet.

Video 3.15.1: Debate" This webcast from March 29, 2011 demonstrates how science is always evolving and how debate among nutrition science experts influences policy decisions.Source: Harvard School of Public Health, in collaboration with Reuters.

Nutritional Science Evolution

One of the newest areas in the realm of nutritional science is the scientific discipline of nutritional genetics, also called nutrigenomics. Genes are part of DNA and contain the genetic information that make up all our traits. Genes are codes for proteins and when they are turned “on” or “off,” they change how the body works. While we know that health is defined as more than just the absence of disease, there are currently very few accurate genetic markers of good health. Rather, there are many more genetic markers for disease. However, science is evolving and nutritional genetics aims to identify what nutrients to eat to “turn on” healthy genes and “turn off” genes that cause disease. Eventually this field will progress so that a person’s diet can be tailored to their genetics. Thus, your DNA will determine your optimal diet.

Video $3.15.2$ Nutrigenomics: A dietitian describes the new scientific discipline of nutrigenomics and its potential impact on diet and health.

Using Science and Technology to Change the Future

As science evolves, so does technology. Both can be used to create a healthy diet, optimize health, and prevent disease. Picture yourself not too far into the future: you are wearing a small “dietary watch” that painlessly samples your blood, and downloads the information to your cell phone, which has an app that evaluates the nutrient profile of your blood and then recommends a snack or dinner menu to assure you maintain adequate nutrient levels. What else is not far off? How about another app that provides a shopping list that adheres to all dietary guidelines and is emailed to the central server at your local grocer who then delivers the food to your home? The food is then stored in your smart fridge which documents your daily diet at home and delivers your weekly dietary assessment to your home computer (Figure $3.15.1$:. At your computer, you can compare your diet with other diets aimed at weight loss, optimal strength training, reduction in risk for specific diseases or any other health goals you may have. You may also delve into the field of nutritional genetics and download your gene expression profiles to a database that analyzes yours against millions of others.

Figure $3.15.1$: The “Smart Fridge". (CC BY 2.0; David Berkowitz)

Key Takeaways

- The scientific method is an organized process of inquiry used in nutritional science to determine if the food suspect fits the claim.
- The scientific method is part of the overall evidence-based approach to designing nutritional guidelines that are based on facts.
- There are different types of scientific studies—epidemiological studies, randomized clinical interventional trial studies, and laboratory animal and cell studies—which all provide different, complementary lines of evidence.

- It takes time to build scientific evidence that culminates as a commonly accepted conclusion.
- Agreement of experts across multiple scientific disciplines is a necessity for recommending dietary changes to improve health and help to prevent disease.
- Science is always evolving as more and more information is collected.

Discussion Starters

1. What are some of the ways in which you think like a scientist and use the scientific method in your everyday life? Any decision-making process uses at least pieces of the scientific method. Think about some of the major decisions you have made in your life and the research you conducted that supported your decision. For example, what computer brand do you own? Where is your money invested? What college do you attend?
2. Do you use technology, appliances, and/or apps that help you to optimize your health?

3.16: Serving Size

Learning Objectives

- Judge food portion sizes for adequacy.

Estimating Portion Size

Have you ever heard the expression, “Your eyes were bigger than your stomach?” This means that you thought you wanted a lot more food than you could actually eat. Amounts of food can be deceiving to the eye, especially if you have nothing to compare them to. It is very easy to heap a pile of mashed potatoes on your plate, particularly if it is a big plate, and not realize that you have just helped yourself to three portions instead of one.

The food industry makes following the *2010 Dietary Guidelines* a challenge. In many restaurants and eating establishments, portion sizes have increased, use of SoFAS has increased, and consequently the typical meal contains more calories than it used to. In addition, our sedentary lives make it difficult to expend enough calories during normal daily activities. In fact, more than one-third of adults are not physically active at all.

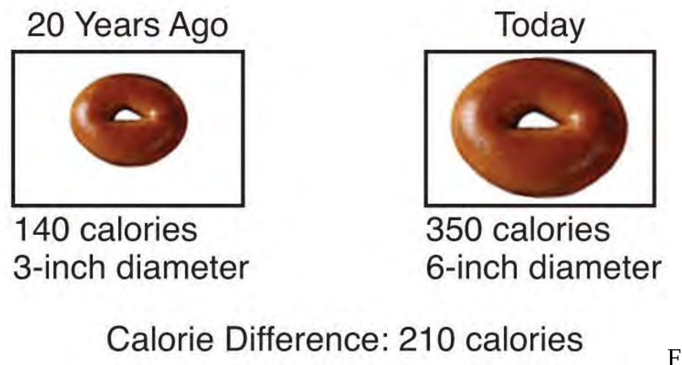


Figure 3.16.1: As food sizes and servings increase it is important to limit the portions of food consumed on a regular basis.

Dietitians have come up with some good hints to help people tell how large a portion of food they really have. Some suggest using common items such as a deck of cards while others advocate using your hand as a measuring rule. See Table 3.16.1 for some examples.American Cancer Society. “Controlling Portion Sizes.” Last revised January 12, 2012. http://www.cancer.org/Healthy/EatHealthyGetActive/TakeControlofYourWeight/controlling-portion-sizes.

Table 3.16.1: Determining Food Portions

Food Product	Amount	Object Comparison	Hand Comparison
Pasta, rice	½ c.	Tennis ball	Cupped hand
Fresh vegetables	1 c.	Baseball	
Cooked vegetables	½ c.	Cupped hand	
Meat, poultry, fish	3 oz.	Deck of cards	Palm of your hand
Milk or other beverages	1 c.	Fist	
Salad dressing	1 Tbsp.	Thumb	
Oil	1 tsp.	Thumb tip	

Video 3.16.1: Managing a Healthy Diet: Judging Healthy Portion Sizes

A dietitian shows how to compare food sizes with hands and other objects.(click to see video)

MyPlate Planner

Estimating portions can be done using the MyPlate Planner. Recall that the MyPlate symbol is divided according to how much of each food group should be included with each meal. Note the MyPlate Planner Methods of Use:

- Fill half of your plate with vegetables such as carrots, broccoli, salad, and fruit.
- Fill one-quarter of your plate with lean meat, chicken, or fish (about 3 ounces)
- Fill one-quarter of your plate with a whole grain such as ⅓ cup rice
- Choose one serving of dairy
- Add margarine or oil for preparation or addition at the table

Table 3.16.2: Meal Planning Guidelines

Carbohydrates	Meats/Proteins	Fats	Free Foods
Choose three servings with each meal.	Choose one to three servings with each meal.	Choose one to two servings with each meal.	Use as desired.
Examples of *one* serving:	Examples of *one* serving:	Examples of *one* serving:	Examples
Breads and Starches • 1 slice bread or small roll • ⅓ c. rice or pasta • ½ c. of cooked cereal or potatoes • ¾ c. dry cereal • ½ c. corn	• 1 oz. lean meat, poultry, or fish • 1 egg • 1 oz. cheese • ¾ c. low-fat cottage cheese	• 1 tsp. margarine, oil, or mayonnaise • 1 Tbsp. salad dressing or cream cheese	**Foods with less than 20 calories per serving.*** • Most vegetables • Sugar-free soda • Black coffee or plain tea
Fruits • 1 piece, such as a small pear • 1 c. fresh fruit • ½ c. canned fruit • ½ c. fruit juice			
Milk • 1 c. skim or low fat • 1 c. unsweetened low-fat yogurt			

Key Takeaways

- Judging portion sizes can be done using your hand or household objects in comparison. It can also be done using the MyPlate guide to determine how much food is a portion for that meal.

Discussion Starter

1. Why is it important to judge portion sizes properly? Explain why it can be tricky to read food labels and figure out the caloric content for one serving.

3.E: Matter and Energy (Exercises)

3.1: In Your Room

3.2: What Is Matter?

1. What is matter?
2. What does weight mean?
3. In this chapter, we'll learn about atoms, which are the building blocks of all matter in the universe. As of 2011, scientists only know of 118 different types of atoms. How do you think it's possible to generate so many different forms of matter using only 118 types of building blocks?
4. Which do you think has more matter, a cup of water or a cup of mercury? Explain.
5. Decide whether each of the following statements is true or false.
    a. Mass and weight are two words for the same concept.
    b. Molecules are bonded together to form atoms.
    c. Alchemists couldn't make gold out of common metals because gold is an element.
    d. The symbol for Gold in the periodic table is Gd.
6. Would you have more mass on the moon or on Earth?
7. Would you have more weight on the moon or on Earth? The force of gravity is stronger on the Earth than it is on the moon.
8. Match the following terms with their meaning.

Terms	**Definitions**
(a) Mass	a. a measure of the total quantity of matter in an object
(b) Volume	b. a measure of how strongly gravity pulls on an object
(c) Weight	c. a measure of the space occupied by an object

9. For the following statements, circle all of the options that apply:
    - Mass depends on…

      (a) the total quantity of matter

      (b) the temperature

      (c) the location

      (d) the force of gravity
    - Volume depends on…

      (a) the total quantity of matter

      (b) the temperature

      (c) the object's shape (independent of size)

      (d) the object's size (independent of shape)
    - Weight depends on…

      (a) the total quantity of matter

      (b) the temperature

      (c) the location

      (d) the force of gravity

3.3: Classifying Matter According to Its State: Solid, Liquid, and Gas

3.4 Classifying Matter According to Its Composition

3.5: Differences in Matter: Physical and Chemical Properties

3.6: Changes in Matter: Physical and Chemical Changes

3.7: Conservation of Mass: There is No New Matter

3.8: Energy

1. Classify each of the following as energy primarily transferred as heat, or energy primarily transferred as work:
   a. The energy transferred from your body to a shopping cart as you push the shopping cart down the aisle.
   b. The energy transferred from a wave to your board when you go surfing.
   c. The energy transferred from the flames to your hotdog when you cook your hotdog over a campfire.
2. Decide whether each of the following statements is true or false:
   a. When heat is transferred to an object, the object cools down.
   b. Any time you raise the temperature of an object, you have done work.
   c. Any time you move an object by applying force, you have done work.
   d. Any time you apply force to an object, you have done work.
3. Rank the following scenarios in order of increasing work:
   a. You apply 100 N of force to a boulder and successfully move it by 2 m.
   b. You apply 100 N of force to a boulder and successfully move it by 1 m.
   c. You apply 200 N of force to a boulder and successfully move it by 2 m.
   d. You apply 200 N of force to a boulder but cannot move the boulder.
4. In science, a vacuum is defined as space that contains absolutely no matter (no molecules, no atoms, etc.) Can energy be transferred as heat through a vacuum? Why or why not?
5. Classify each of the following energies as kinetic energy or potential energy:
   a. The energy in a chocolate bar.
   b. The energy of rushing water used to turn a turbine or a water wheel.
   c. The energy of a skater gliding on the ice.
   d. The energy in a stretched rubber band.
6. Decide which of the following objects has more kinetic energy:
   a. A 200 lb. man running at 6 mph or a 200 lb. man running at 3 mph.
   b. A 200 lb. man running at 7 mph or a 150 lb. man running at 7 mph.
   c. A 400 lb. man running at 5 mph or a 150 lb. man running at 3 mph.
7. A car and a truck are traveling along the highway at the same speed.
   a. If the car weighs 1500 kg and the truck weighs 2500 kg, which has more kinetic energy, the car or the truck?
   b. Both the car and the truck convert the potential energy stored in gasoline into the kinetic energy of motion. Which do you think uses more gas to travel the same distance, the car or the truck?
8. You mix two chemicals in a beaker and notice that as the chemicals react, the beaker becomes noticeably colder. Which chemicals have more chemical potential energy, those present at the start of the reaction or those present at the end of the reaction?

3.9: Energy and Chemical and Physical Change

3.10: Temperature: Random Motion of Molecules and Atoms

3.11: Temperature Changes: Heat Capacity

3.12: Energy and Heat Capacity Calculations

1. A pot of water is set on a hot burner of a stove. What is the direction of heat flow?
2. Some uncooked macaroni is added to a pot of boiling water. What is the direction of heat flow?
3. How much energy in calories is required to heat 150 g of H_2O from 0°C to 100°C?
4. How much energy in calories is required to heat 125 g of Fe from 25°C to 150°C?
5. If 250 cal of heat were added to 43.8 g of Al at 22.5°C, what is the final temperature of the aluminum?

6. If 195 cal of heat were added to 33.2 g of Hg at 56.2°C, what is the final temperature of the mercury?
7. A sample of copper absorbs 145 cal of energy, and its temperature rises from 37.8°C to 41.7°C. What is the mass of the copper?
8. A large, single crystal of sodium chloride absorbs 98.0 cal of heat. If its temperature rises from 22.0°C to 29.7°C, what is the mass of the NaCl crystal?
9. If 1.00 g of each substance in Table 7.3 were to absorb 100 cal of heat, which substance would experience the largest temperature change?
10. If 1.00 g of each substance in Table 7.3 were to absorb 100 cal of heat, which substance would experience the smallest temperature change?
11. Determine the heat capacity of a substance if 23.6 g of the substance gives off 199 cal of heat when its temperature changes from 37.9°C to 20.9°C.
12. What is the heat capacity of gold if a 250 g sample needs 133 cal of energy to increase its temperature from 23.0°C to 40.1°C?

Answers

1. Heat flows into the pot of water.
3. 15,000 cal
5. 49.0°C
7. 404 g
9. Mercury would experience the largest temperature change.
11. 0.496 cal/g•°C

3.E: Nutrition and You (Exercises)

Study Questions:

1. What does Essential mean?

2. What are the 6 classes of essential nutrients? Put a star by the MOST essential nutrient.

3. How many of the 6 classes of essential nutrients are present in the food you eat?

4. Which of the nutrients give you Calories?

5. If you ate a food that contained: 10 grams of Carbohydrate, 10 grams of Fat and 10 grams of Protein, how many Calories would that food contain (approximately)? *Hint: Use 4 Calories per gram of Carbohydrate, 9 Calories per gram of Fat and 4 Calories per gram of Protein*

6. Is Alcohol a nutrient? Does Alcohol give you Calories?

7. When it comes to nutrients, is it “the more the better”? Why or Why not?

8. Are there “good foods” and “bad foods”? Why or Why not?

9. What are “phytochemicals”?

10. What is meant by the term "serving size"? *For example, a "serving size" of potato chips is 1 ounce or about 15 chips.*

11. What is your opinion of the role of the federal government in promoting health and preventing disease in Americans?

Expand Your Knowledge

1. Conduct an assessment of your diet. Begin by recording what you eat every day, including snacks and beverages, in a journal. Then visit the website, myfitnesspal.com which has tools to help you assess your diet and exercise. Calculate how many Calories you are eating in a day compared to how many are suggested for you per day.
2. Conduct an informal survey of five of your friends and family members. Create a questionnaire with ten to fifteen questions pertaining to their perception of food, their favorite food shows, what restaurants they frequent, what they purchase from the grocery store each week, etc. Ask questions about their general eating habits and record the answers. Next, get a camera and take pictures of the insides of their refrigerators. What have you learned about their eating habits? What advice would you give them? Review all the material and put it together in a report. Be prepared to share your findings.
3. Define eight steps your community can take to build a more sustainable food system. Visit the websites of Toronto’s and California’s plans to help provide you with some good ideas to accomplish the challenge. Toronto’s plan for a sustainable food system: wx.toronto.ca/inter/health/fo...s%20report.pdf

   California’s plan for a sustainable food system: www.vividpicture.net/

CHAPTER OVERVIEW

4: Atoms and Elements

4.1: Cutting Aluminum until you get Atoms

Take some aluminum foil. Cut it in half. Now there are two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil. It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can this exercise be taken, at least in theory? Can one continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC by Democritus and other ancient Greek philosophers (Figure 4.1.1).

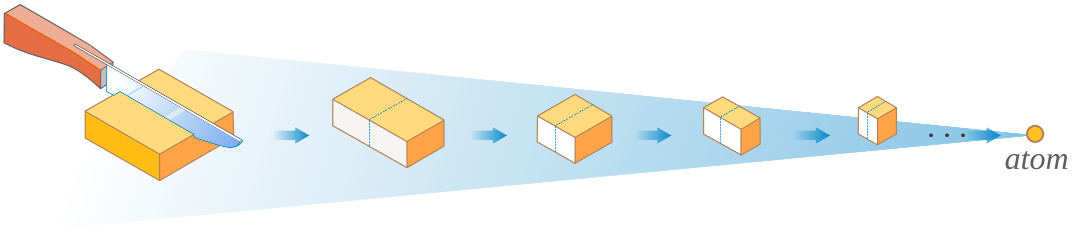


Figure 4.1.1: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya via LibreTexts)

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called diatomic molecules. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

4.2: Indivisible - The Atomic Theory

Learning Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

You learned earlier that all matter in the universe is made out of tiny building blocks called atoms. All modern scientists accept the concept of the atom, but when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus and the Greek Philosophers

About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge, entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable "thing." One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon—tiny, indivisible, solid objects—make up all matter in the universe.

Figure 4.2.1: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

Democritus then reasoned that changes occur when the many atomos in an object were reconnected or recombined in different ways. Democritus even extended this theory, suggesting that there were different varieties of atomos with different shapes, sizes, and masses. He thought, however, that shape, size, and mass were the only properties differentiating the different types of atoms. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves, but rather, resulted from the different ways in which the atomos were combined and connected to one another.

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they are known today) was fully appreciated.

Greeks: "All Thought and No Action"

Greek philosophers were "all thought and no action" and did not feel the need to test their theories with reality. In contrast, Dalton's efforts were based on experimentation and testing ideas against reality.

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

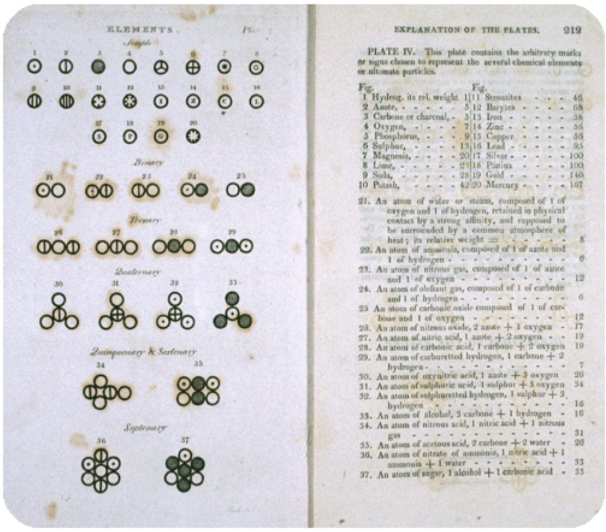
The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 4.2.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory were as follows:

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.



EXPLANATION OF THE PLATES. 219

PLATE IV. This plate contains the arbitrary marks or signs chosen to represent the several chemical elements or ultimate particles.

Fig.			Fig.		
1	Hydrog. its rel. weight	1	11	Strontites	46
2	Azote,	5	12	Barytes	68
3	Carbone or charcoal,	5	13	Iron	38
4	Oxygen,	7	14	Zinc	56
5	Phosphorus,	9	15	Copper	56
6	Sulphur,	13	16	Lead	95
7	Magnesia,	20	17	Silver	100
8	Lime,	24	18	Platina	100
9	Soda,	28	19	Gold	140
10	Potash,	42	20	Mercury	167

21. An atom of water or steam, composed of 1 of oxygen and 1 of hydrogen, retained in physical contact by a strong affinity, and supposed to be surrounded by a common atmosphere of heat; its relative weight = - - - - - - 8
22. An atom of ammonia, composed of 1 of azote and 1 of hydrogen - - - - - - - - - - - 6
23. An atom of nitrous gas, composed of 1 of azote and 1 of oxygen - - - - - - - - - - 12
24. An atom of olefiant gas, composed of 1 of carbone and 1 of hydrogen - - - - - - - - - - 6
25 An atom of carbonic oxide composed of 1 of carbone and 1 of oxygen - - - - - - - - - 12
26. An atom of nitrous oxide, 2 azote + 1 oxygen - 17
27. An atom of nitric acid, 1 azote + 2 oxygen - - 19
28. An atom of carbonic acid, 1 carbone + 2 oxygen 19
29. An atom of carburetted hydrogen, 1 carbone + 2 hydrogen - - - - - - - - - - - - - 7
30. An atom of oxynitric acid, 1 azote + 3 oxygen 26
31. An atom of sulphuric acid, 1 sulphur + 3 oxygen 34
32. An atom of sulphuretted hydrogen, 1 sulphur + 3 hydrogen - - - - - - - - - - - - - 16
33. An atom of alcohol, 3 carbone + 1 hydrogen - 16
34. An atom of nitrous acid, 1 nitric acid + 1 nitrous gas - - - - - - - - - - - - - - 31
35. An atom of acetous acid, 2 carbone + 2 water - 26
36. An atom of nitrate of ammonia, 1 nitric acid + 1 ammonia + 1 water - - - - - - - - 33
37. An atom of sugar, 1 alcohol + 1 carbonic acid - 35

Figure 4.2.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video 4.2.1, control of individual atoms can be use used create animations.

Video 4.2.1: A Boy And His Atom: The World's Smallest Movie.

A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos." However, other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.
- Dalton's Atomic Theory is the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom. The general tenets of this theory are:
  - All matter is composed of extremely small particles called atoms.
  - Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
  - Atoms cannot be subdivided, created, or destroyed.
  - Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
  - In chemical reactions, atoms are combined, separated, or rearranged.

4.3: The Nuclear Atom

Learning Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" model of the atom and the evidence for it.
- Draw a diagram of Thomson's "plum pudding" model of the atom and explain why it has this name.
- Describe Rutherford's gold foil experiment and explain how this experiment altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.

Figure 4.3.1: J. J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom.

Electrons and Plums

The electron was discovered by J. J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This was a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or similar to a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges, and the electron would be pulled back more toward the center region of the atom.

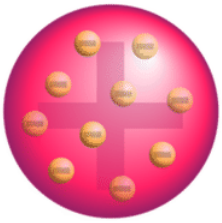


Figure 4.3.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871-1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Atoms and Gold

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

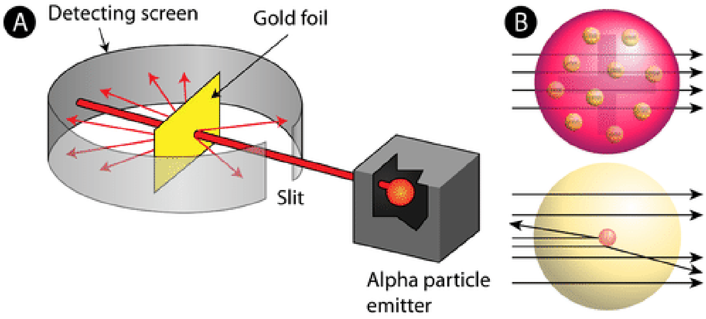


Figure 4.3.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupy the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Atomic Nucleus

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. The figure below shows these parts of the atom.

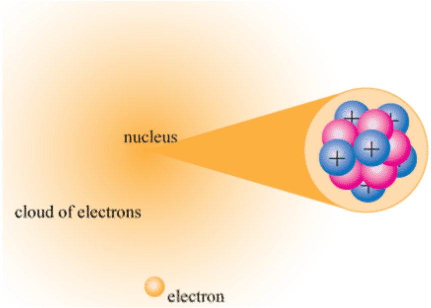


Figure 4.3.4: The nuclear atom.

The nucleus of the atom is extremely small. Its radius is only about 1/100,000 of the total radius of the atom. Electrons have virtually no mass, but protons and neutrons have a lot of mass for their size. As a result, the nucleus has virtually all the mass of an atom. Given its great mass and tiny size, the nucleus is very dense. If an object the size of a penny had the same density as the nucleus of an atom, its mass would be greater than 30 million tons!

Holding it all Together

Particles with opposite electric charges attract each other. This explains why negative electrons orbit the positive nucleus. Particles with the same electric charge repel each other. This means that the positive protons in the nucleus push apart from one another. So why doesn't the nucleus fly apart? An even stronger force—called the **strong nuclear force**—holds protons and neutrons together in the nucleus.

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

4.4: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it was not entirely accurate, because contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to us: protons and neutrons. We already learned that J. J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus, called a **proton**. There is also a third subatomic particle, known as a neutron.

Electrons

Electrons are one of three main types of particles that make up atoms. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particle called leptons. All leptons have an electric charge of -1 or 0. Electrons are extremely small. The mass of an electron is only about 1/2000 the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of -1, which is equal but opposite to the charge of a proton, which is $+1$. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract one another, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. However, this is an incorrect perspective, as quantum mechanics demonstrates that electrons are more complicated.

Figure 4.4.1: Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

A proton is one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one $(+1)$ and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom.

Neutrons

Atoms of all elements—except for most atoms of hydrogen—have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge—they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they are bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 4.4.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	−1	outside the nucleus
neutron	n^0	1	1	0	inside the nucleus

Table 4.4.1 gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they are neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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4.5: Elements- Defined by Their Number of Protons

Learning Objectives

- Define atomic number.
- Define mass number.
- Determine the number of protons, neutrons, and electrons in an atom.

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as H for hydrogen, C for carbon, or He for helium. However, it would more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.

Figure 4.5.1: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 4.5.1). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (Z). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 4.5.1: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, 79 electrons as well.

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table 4.5.1). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table 4.5.1).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units ($2\ \mathrm{amu}$ for the protons + $2\ \mathrm{amu}$ for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called isotopes and are discussed in detail later on).

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

An atom's mass number is very easy to calculate, provided that you know the number of protons and neutrons in an atom.

Example 4.5.1

What is the mass number of an atom of helium that contains 2 neutrons?

Solution

$(\text{number of protons}) = 2$ (Remember that an atom of helium always has 2 protons.)

$(\text{number of neutrons}) = 2$

$\text{mass number} = (\text{number of protons}) + (\text{number of neutrons})$

$$\text{mass number} = \mathbf{2 + 2 = 4}$$

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are O for oxygen, Zn for zinc, and Fe for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table 4.5.2

Table 4.5.2: *Symbols and Latin Names for Elements*

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum

Chemical Symbol	Name	Latin Name
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

4.6: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge as a result because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called **cations**. Most metals become cations when they make ionic compounds.

Cations

A neutral sodium atom is likely to achieve an octet in its outermost shell by losing its one valence electron.

$$\mathrm{Na} \rightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure $4.6.1$ is a graphical depiction of this process.

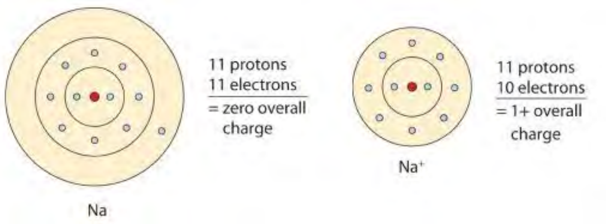

Figure $4.6.1$: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called **anions**. Most nonmetals become anions when they make ionic compounds.

A neutral chlorine atom has seven electrons in its outermost shell. Only one more electron is needed to achieve an octet in chlorine's valence shell. (In table salt, this electron comes from the sodium atom.)

$$\mathrm{e}^{-} + \mathrm{Cl} \longrightarrow \mathrm{Cl}^{-}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (-*ide* instead of -*ine*) to create the name of this anion. Figure $4.6.2$ is a graphical depiction of this process.

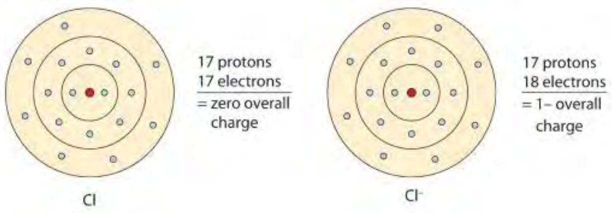

Figure 4.6.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge.

Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall charge.

> *The names for positive and negative ions are pronounced CAT-eye-ons and ANN-eye-ons, respectively.*

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge. Figure 4.6.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written Ba^{2+}, not Ba^{+2}.

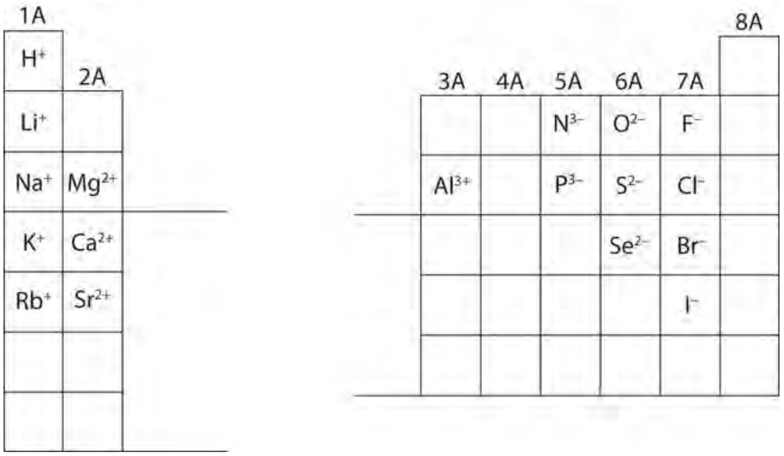

Figure 4.6.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

Contributions & Attributions

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4.7: Isotopes - When the Number of Neutrons Varies

Learning Objectives

- Explain what isotopes are and how an isotope affects an element's atomic mass.
- Determine the number of protons, electrons, and neutrons of an element with a given mass number.

All atoms of the same element have the same number of protons, but some may have different numbers of neutrons. For example, all carbon atoms have six protons, and most have six neutrons as well. But some carbon atoms have seven or eight neutrons instead of the usual six. Atoms of the same element that differ in their numbers of neutrons are called **isotopes**. Many isotopes occur naturally. Usually one or two isotopes of an element are the most stable and common. Different isotopes of an element generally have the same physical and chemical properties because they have the same numbers of protons and electrons.

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 4.7.1. Most hydrogen atoms have just one proton, one electron, and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

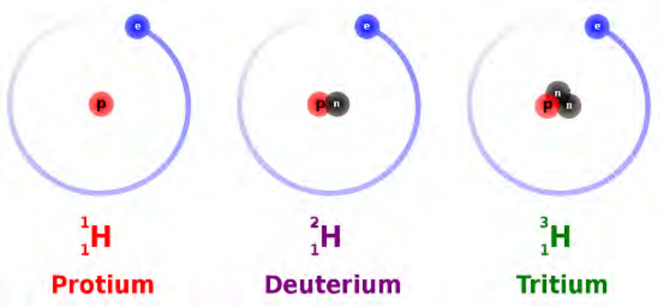


Figure 4.7.1: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

For most elements other than hydrogen, isotopes are named for their mass number. For example, carbon atoms with the usual 6 neutrons have a mass number of 12 (6 protons + 6 neutrons = 12), so they are called carbon-12. Carbon atoms with 7 neutrons have an atomic mass of 13 (6 protons + 7 neutrons = 13). These atoms are the isotope called carbon-13.

Example 4.7.1: Lithium Isotopes

a. What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons?
b. What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons?

Solution

A lithium atom contains 3 protons in its nucleus irrespective of the number of neutrons or electrons.

a.

$$\begin{aligned}\text{atomic number} = (\text{number of protons}) &= 3\\ (\text{number of neutrons}) &= 3\end{aligned}$$

$$\begin{aligned}\text{mass number } &= (\text{number of protons}) + (\text{number of neutrons})\\ \text{mass number } &= 3+3\\ &= 6\end{aligned}$$

b.

$$\begin{aligned}\text{atomic number} = (\text{number of protons}) &= 3\\ (\text{number of neutrons}) &= 4\end{aligned}$$

$$\begin{aligned}\text{mass number } &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number } &= 3 + 4 \\ &= 7\end{aligned}$$

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons, and as an isotope with 4 neutrons, but it doesn't exist as an isotope with 2 neutrons or as an isotope with 5 neutrons.

Stability of Isotopes

Atoms need a certain ratio of neutrons to protons to have a stable nucleus. Having too many or too few neutrons relative to protons results in an unstable, or radioactive, nucleus that will sooner or later break down to a more stable form. This process is called radioactive decay. Many isotopes have radioactive nuclei, and these isotopes are referred to as radioisotopes. When they decay, they release particles that may be harmful. This is why radioactive isotopes are dangerous and why working with them requires special suits for protection. The isotope of carbon known as carbon-14 is an example of a radioisotope. In contrast, the carbon isotopes called carbon-12 and carbon-13 are stable.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundance"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element—chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below:

Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

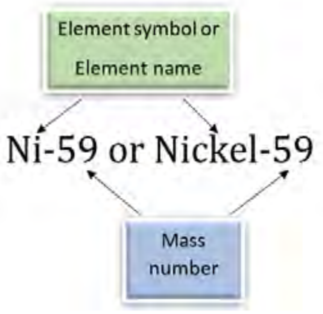
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\text{Ni}$$

$$^{238}_{92}\text{U}$$

In the nickel nucleus represented above, the atomic number 28 indicates that the nucleus contains 28 protons, and therefore, it must contain 31 neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons, as all uranium nuclei do; and this particular uranium nucleus has 146 neutrons.

Another way of representing isotopes is by adding a hyphen and the mass number to the chemical name or symbol. Thus the two nuclei would be Nickel-59 or Ni-59 and Uranium-238 or U-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) are given to the side of the name.



✓ Example 4.7.2: Potassium-40

How many protons, electrons, and neutrons are in an atom of ${}^{40}_{19}\text{K}$?

Solution

$$\text{atomic number} = (\text{number of protons}) = 19$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 19$$

The mass number, 40, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 40 - 19 = 21.$$

✓ Example 4.7.3: Zinc-65

How many protons, electrons, and neutrons are in an atom of zinc-65?

Solution

$$\text{number of protons} = 30$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 30$$

The mass number, 65, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 65 - 30 = 35$$

? Exercise 4.7.3

How many protons, electrons, and neutrons are in each atom?

a. ${}^{60}_{27}\text{Co}$
b. Na-24
c. ${}^{45}_{20}\text{Ca}$
d. Sr-90

Answer a:

27 protons, 27 electrons, 33 neutrons

Answer b:

11 protons, 11 electrons, 13 neutrons

Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Summary

- The number of protons is always the same in atoms of the same element.
- The number of neutrons can be different, even in atoms of the same element.
- Atoms of the same element that contain the same number of protons, but different numbers of neutrons, are known as **isotopes**.
- Isotopes of any given element all contain the same number of protons, so they have the same atomic number (for example, the atomic number of helium is always 2).
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

4.7: Isotopes - When the Number of Neutrons Varies is shared under a CK-12 license and was authored, remixed, and/or curated by Marisa Alviar-Agnew & Henry Agnew.

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4.8: Atomic Mass - The Average Mass of an Element’s Atoms

Learning Objectives

- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units.

Calculating Atomic Mass

You can calculate the atomic mass (or average mass) of an element provided you know the **relative abundance** (the fraction of an element that is a given isotope), the element's naturally occurring isotopes, and the masses of those different isotopes. We can calculate this by the following equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

Look carefully to see how this equation is used in the following examples.

Example 4.8.1: Boron Isotopes

Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and mass of 10 amu. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of 11 amu. What is the atomic mass of boron?

Solution

Boron has two isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.20$ (Write all percentages as decimals), $\text{mass}_1 = 10$
- Isotope 2: $\%_2 = 0.80$, $\text{mass}_2 = 11$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)(10) + (0.80)(11)$$

$$\text{Atomic mass} = 10.8\text{ amu}$$

The mass of an average boron atom, and thus boron's atomic mass, is 10.8 amu

Example 4.8.2: Neon Isotopes

Neon has three naturally occurring isotopes. In a sample of neon, 90.92% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of 19.99 amu Another 0.3% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of 20.99 amu The final 8.85% of the atoms are Ne-22, which is an isotope of neon with 12 neutrons and a mass of 21.99 amu What is the atomic mass of neon?

Solution

Neon has three isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.9092$ (write all percentages as decimals), $\text{mass}_1 = 19.99$
- Isotope 2: $\%_2 = 0.003$, $\text{mass}_2 = 20.99$
- Isotope 3: $\%_3 = 0.0885$, $\text{mass}_3 = 21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99)$$

$$\text{Atomic mass} = 20.17 \text{ amu}$$

The mass of an average neon atom is $\mathbf{20.17}$ amu

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol B) is 10.8, which is what we calculated in Example $4.8.1$, and the atomic mass of neon (symbol Ne) is 20.8, which is what we calculated in Example $4.8.2$. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

? Exercise 4.8.1

Chlorine has two naturally occurring isotopes. In a sample of chlorine, 75.77% of the atoms are Cl-35, with a mass of 34.97 amu Another 24.23% of the atoms are Cl-37, with a mass of 36.97 amu What is the atomic mass of chlorine?

Answer

35.45 amu

Summary

- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- An element's atomic mass can be calculated provided the relative abundance of the element's naturally occurring isotopes and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.

4.9: Arrangements of Electrons

Learning Objectives

- Describe how electrons are grouped within atoms.

Although we have discussed the general arrangement of subatomic particles in atoms, we have said little about how electrons occupy the space about the nucleus. Do they move around the nucleus at random, or do they exist in some ordered arrangement?

The modern theory of electron behavior is called quantum mechanics. It makes the following statements about electrons in atoms:

- Electrons in atoms can have only certain specific energies. We say that the energies of the electrons are quantized.
- Electrons are organized according to their energies into sets called **shells** (labeled by the principle quantum number, ***n***). Generally the higher the energy of a shell, the farther it is (on average) from the nucleus. Shells do not have specific, fixed distances from the nucleus, but an electron in a higher-energy shell will spend more time farther from the nucleus than does an electron in a lower-energy shell.
- Shells are further divided into subsets of electrons called **subshells**. The first shell has only one subshell, the second shell has two subshells, the third shell has three subshells, and so on. The subshells of each shell are labeled, in order, with the letters *s*, *p*, *d*, and *f*. Thus, the first shell has only a single *s* subshell (called **1*s***), the second shell has **2*s*** and **2*p*** subshells, the third shell has **3*s***, **3*p***, and **3*d*** and so forth.

Table 4.9.1: Shells and Subshells

Shell	Number of Subshells	Names of Subshells
1	1	*1s*
2	2	*2s* and *2p*
3	3	*3s*, *3p* and *3d*
4	4	*4s*, *4p*, *4d* and *4f*

- Different subshells hold a different maximum number of electrons. **Any *s* subshell can hold up to 2 electrons; *p*, 6; *d*, 10; and *f*, 14.**

Table 4.9.2: Number of Electrons

Subshell	Maximum Number of Electrons
s	2
p	6
d	10
f	14

It is the arrangement of electrons into shells and subshells that most concerns us here, so we will focus on that.

We use numbers to indicate which shell an electron is in. As shown in Table 4.9.1, the first shell, closest to the nucleus and with the lowest-energy electrons, is shell 1. This first shell has only one subshell, which is labeled 1*s* and can hold a maximum of 2 electrons. We combine the shell and subshell labels when referring to the organization of electrons about a nucleus and use a superscript to indicate how many electrons are in a subshell. Thus, because a hydrogen atom has its single electron in the *s* subshell of the first shell, we use $\mathbf{1}\boldsymbol{s}^{\mathbf{1}}$ to describe the electronic structure of hydrogen. This structure is called an electron configuration. Electron configurations are shorthand descriptions of the arrangements of electrons in atoms. The electron configuration of a hydrogen atom is spoken out loud as “one-ess-one.”

Helium atoms have 2 electrons. Both electrons fit into the 1*s* subshell because *s* subshells can hold up to 2 electrons; therefore, the electron configuration for helium atoms is $\mathbf{1}\boldsymbol{s}^{\mathbf{2}}$ (spoken as “one-ess-two”).

The 1*s* subshell cannot hold 3 electrons (because an *s* subshell can hold a maximum of 2 electrons), so the electron configuration for a lithium atom cannot be $1s^3$. Two of the lithium electrons can fit into the 1*s* subshell, but the third electron must go into the second shell. The second shell has two subshells, *s* and *p*, which fill with electrons in that order. The 2*s* subshell holds a maximum of 2 electrons, and the 2*p* subshell holds a maximum of 6 electrons. Because lithium's final electron goes into the 2*s* subshell, we write the electron configuration of a lithium atom as **$1s^2 2s^1$.** The shell diagram for a lithium atom is shown below. The shell closest to the nucleus (first shell) has 2 dots representing the 2 electrons in **1*s***, while the outermost shell (**2*s***) has 1 electron.

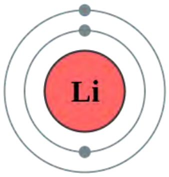


Figure 4.9.1: Shell diagram of lithium (Li) atom.

The next largest atom, beryllium, has 4 electrons, so its electron configuration is $1s^2 2s^2$. Now that the 2*s* subshell is filled, electrons in larger atoms start filling the 2*p* subshell. Thus, the electron configurations for the next six atoms are as follows:

- B: $1s^2 2s^2 2p^1$
- C: $1s^2 2s^2 2p^2$
- N: $1s^2 2s^2 2p^3$
- O: $1s^2 2s^2 2p^4$
- F: $1s^2 2s^2 2p^5$
- Ne: $1s^2 2s^2 2p^6$

With neon, the 2*p* subshell is completely filled. Because the second shell has only two subshells, atoms with more electrons now must begin the third shell. The third shell has three subshells, labeled *s*, *p*, and *d*. The *d* subshell can hold a maximum of 10 electrons. The first two subshells of the third shell are filled in order—for example, the electron configuration of aluminum, with 13 electrons, is $1s^2 2s^2 2p^6 3s^2 3p^1$. However, a curious thing happens after the 3*p* subshell is filled: the 4*s* subshell begins to fill before the 3*d* subshell does. In fact, the exact ordering of subshells becomes more complicated at this point (after argon, with its 18 electrons), so we will not consider the electron configurations of larger atoms. A fourth subshell, the *f* subshell, is needed to complete the electron configurations for all elements. An *f* subshell can hold up to 14 electrons.

Electron filling always starts with **1*s***, the subshell closest to the nucleus. Next is **2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d***, **4*p***, **5s, 4d, 5p, 6s,** etc., shown in the electron shell filling order diagram in Figure 4.9.2. Follow each arrow in order **from top to bottom**. The subshells you reach along each arrow give the ordering of filling of subshells in larger atoms.

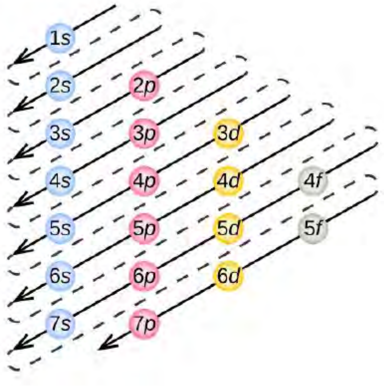


Figure 4.9.2: The order of electron filling in an atom.

✓ Example 4.9.1: Electronic Configuration of Phosphorus Atoms

Using Figure 4.9.2 as your guide, write the electron configuration of a neutral phosphorus atom. The atomic number of P is 15.

Solution

A neutral phosphorus atom has 15 electrons. Two electrons can go into the 1*s* subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 5 electrons. Of those 5 electrons, 2 can go into the 3*s* subshell, and the remaining 3 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral phosphorus atoms is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Exercise 4.9.1: Electronic Configuration of Chlorine Atoms

Using Figure 4.9.2 as your guide, write the electron configuration of a neutral chlorine atom. The atomic number of Cl is 17.

Answer

A neutral chlorine atom has 17 electrons. Two electrons can go into the 1*s* subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 7 electrons. Of those 7 electrons, 2 can go into the 3*s* subshell, and the remaining 5 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral chlorine atoms is $1s^22s^22p^63s^23p^5$.

Since the arrangement of the periodic table is based on the electron configurations, Figure 4.9.3 provides an alternative method for determining the electron configuration. The filling order simply **begins at the top left,** with hydrogen (Z=1) and includes each subshell as you proceed in increasing atomic number (*Z)* order.

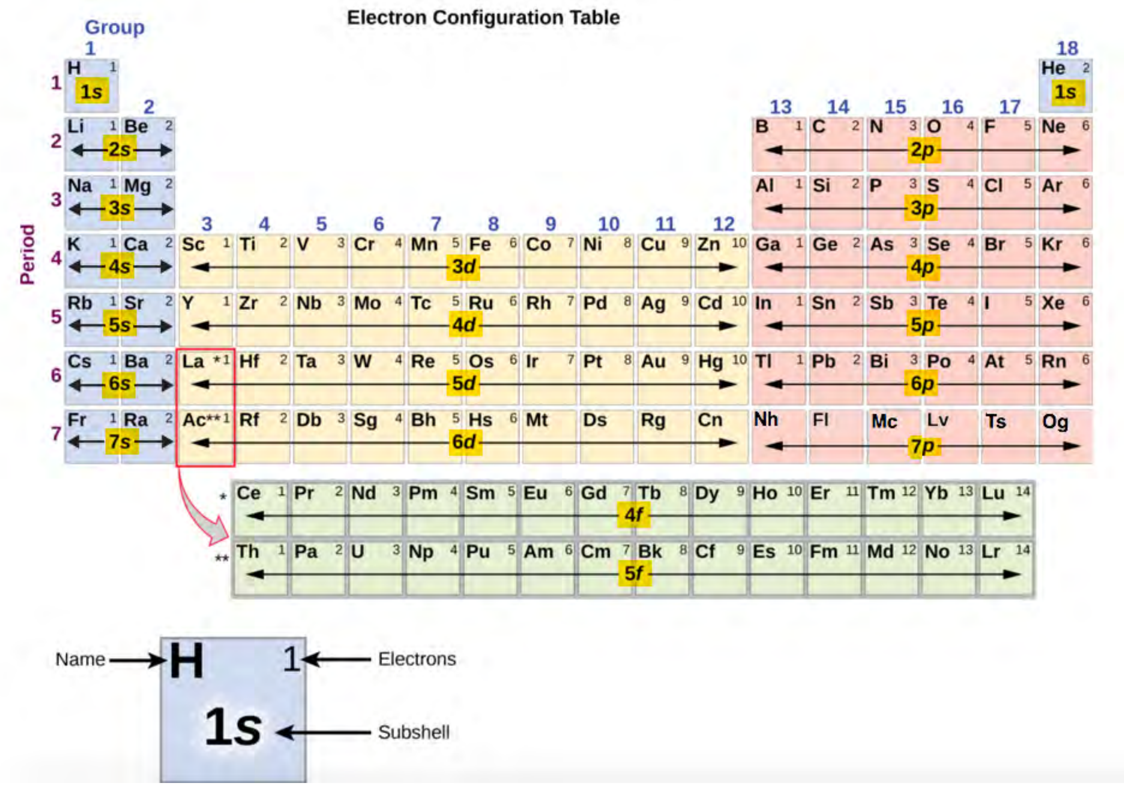


Figure 4.9.3: This periodic table shows the electron configuration for each subshell. By “building up” from hydrogen, this table can be used to determine the electron configuration for any atom on the periodic table.

For example, the first row (Period 1) contains H and He only, because only two electrons are required to fill the 1s subshell. The second row s-block, contains only two elements, Li and Be, to fill the 2s subshell. This is followed by the second row p-block, containing 6 elements (B through Ne) since six electrons are required to fill the 2p subshell. The third row is similar to the second row elements. Two electrons are needed (Na and Mg) to fill the 3s subshell and six electrons are required (Al through Ar) to complete the 3p subshell. After filling the 3*p* block up to Ar, we see the next subshell will be 4s (K, Ca), followed by the 3*d* subshell, which are filled by ten electrons (Sc through Zn). The 4p subshell is filled next by six electrons (Ga through Kr). As you can see, the periodic table shown in Figure 4.9.3 provides a simple way to remember the order of filling the subshells in determining the electron configuration. The order of filling subshells is the same: **1s, 2*s*, 2*p*, 3s, 3*p*, 4*s*, 3*d***, **4*p***, **5s, 4d, 5p, 6s,** etc.

Example 4.9.2: Aluminum

Using Figure 4.9.3 as your guide, write the electron configuration of neutral aluminum atom. The atomic number of Al is 13.

Solution

Aluminum has 13 electrons.

Start at Period 1 of the periodic table, Figure 4.9.3. Place two electrons in the 1s subshell (**$1s^2$**).

Proceed to Period 2 (left to right direction). Place the next two electrons in the 2s subshell (**$2s^2$**) and the next six electrons in the 2p subshell (**$2p^6$**).

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell (**$3s^2$**) and the last one electron in the 3p subshell (**$3p^1$**).

The electron configuration of Aluminum is **$1s^22s^22p^63s^23p^1$**

rcise

Using Figure 4.9.3 as your guide, write the electron configuration of the atom that has 20 electrons

Answer

Start at Period 1 of Figure 4.9.3 Place two electrons in the 1s subshell (**$1s^2$**).

Proceed to Period 2 (left to right direction). Place the next two electrons in the 2s subshell (**$2s^2$**) and the next six electrons in the 2p subshell (**$2p^6$**).

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell (**$3s^2$**) and the next six electron in the 3p subshell (**$3p^6$**).

Proceed to Period 4. Place the remaining two electrons in the 4s subshell (**$4s^2$).**

The electron configuration is **$1s^22s^22p^63s^23p^64s^2$**

Valence Electrons

In the study of chemical reactivity, we will find that the electrons in the outermost principal energy level are very important and so they are given a special name. **Valence electrons** are the electrons in the highest occupied principal energy level of an atom.

In the second period elements, the two electrons in the $1s$ sublevel are called **inner-shell electrons** and are not involved directly in the element's reactivity or in the formation of compounds. Lithium has a single electron in the second principal energy level and so we say that lithium has one valence electron. Beryllium has two valence electrons. How many valence electrons does boron have? You must recognize that the second principal energy level consists of both the $2s$ and the $2p$ sublevels and so the answer is three. In fact, the number of valence electrons goes up by one for each step across a period until the last element is reached. Neon, with its configuration ending in $2s^22p^6$, has eight valence electrons.

The alkali metal sodium (atomic number 11) has one more electron than the neon atom. This electron must go into the lowest-energy subshell available, the 3*s* orbital, giving a $1s^22s^22p^63s^1$ configuration. The electrons occupying the outermost shell orbital(s) (highest value of *n*) are called valence electrons, and those occupying the inner shell orbitals are called **core electrons** (Figure \PageIndex4). Since the core electron shells correspond to noble gas electron configurations, we can abbreviate electron configurations by writing the noble gas that matches the core electron configuration, along with the valence electrons in a condensed format. For our sodium example, the symbol [Ne] represents core electrons, ($1s^22s^22p^6$) and our abbreviated or condensed configuration is [Ne]$3s^1$.

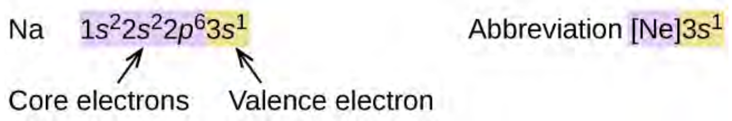


Figure 4.9.4: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.

Similarly, the abbreviated configuration of lithium can be represented as [He]$2s^1$, where [He] represents the configuration of the helium atom, which is identical to that of the filled inner shell of lithium. Writing the configurations in this way emphasizes the similarity of the configurations of lithium and sodium. Both atoms, which are in the alkali metal family, have only one electron in a valence *s* subshell outside a filled set of inner shells.

$$\text{Li}: [\text{He}]\, 2s^1$$
$$\text{Na}: [\text{Ne}]\, 3s^1$$

A chemical reaction results from electron removal, electron addition, or electron sharing of the valence electrons of the different atoms. The path a specific element will take depends on where the electrons are in the atom and how many there are. Thus, it is convenient to separate electrons into two groups. Valence shell electrons (or, more simply, the *valence electrons*) are the electrons in the highest-numbered shell, or valence shell, while core electrons are the electrons in lower-numbered shells. We can see from the electron configuration of a carbon atom—$1s^22s^22p^2$—that it has 4 valence electrons ($2s^22p^2$) and 2 core electrons ($1s^2$). You will see in the next chapters that the chemical properties of elements are determined by the number of valence electrons.

✓ Example 4.9.3

Examine the electron configuration of neutral phosphorus atoms in Example 4.9.1, $1s^22s^22p^63s^23p^3$ and write the abbreviated notation.

Solution

Phosphorus has electron configuration, $1s^22s^22p^63s^23p^3$.

The highest-numbered shell is the third shell ($3s^23p^3$): **2 electrons** in the 3*s* subshell and **3 electrons** in the 3*p* subshell. That gives a total of **5 valence electrons**.

The 10 inner shell (core) electrons, $1s^22s^22p^6$ can be replaced by [Ne] (see Figure 4.9.3). Abbreviated notation is : [Ne]$3s^23p^3$

? Exercise 4.9.3

Examine the electron configuration of neutral calcium atom (Exercise 4.9.2), $1s^22s^22p^63s^23p^64s^2$, and write the abbreviated notation.

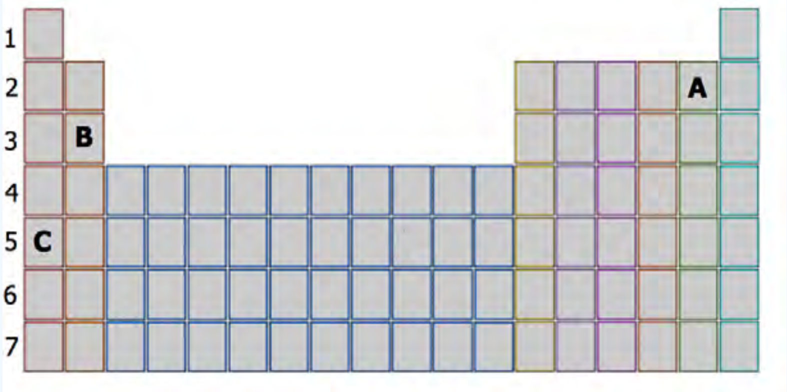
Answer

The highest-numbered shell is the fourth shell $4s^2$, which has **2 electrons** in the 4*s* subshell. Hence, Calcium has **2 valence electrons**.

The 18 inner-shell (core) electrons, $1s^22s^22p^63s^23p^6$, can be replaced by [Ar], see Figure 4.9.3 The abbreviated notation is: [Ar]$4s^2$

✓ Example 4.9.4

Based on their respective locations in the periodic table (use Figure 4.9.3), determine the number of valence electrons and the valence shell configuration of elements A, B and C.



Solution

Element A is located in Period 2, the *5th position* in *2p*-block. Before the electrons are placed in *2p* subshell, the *2s* subshell must be filled first. This means that A has ***two valence electrons*** in 2s (**$2s^2$**) and ***five valence electrons*** in 2p (**$2p^5$**). Answer: **$2s^2 2p^5$.** It has ***2 + 5 = 7* valence electrons.**

Element B is located in Period 3, the *2nd position* in *3s*-block. This means that B has ***two valence electrons*** in 3s (**$3s^2$**). Answer: **$3s^2$.**

Element C is located in Period 5, the *1st position* in *5s*-block). This means that there is only **one valence electron** in 5s (**$5s^1$**). Answer: **$5s^1$**.

rcise

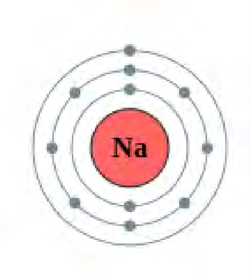
Using the location of Na is the periodic table (Figure 4.9.3), draw the shell diagram of sodium atom.

Answer

Sodium (Na) is the first element in the 3rd row (Period 3) in the periodic table. This means that the first shell and second shells of Na atom are filled to the maximum number of electrons.

The first shell **(1s)** *is filled with **2 electrons***. The second shell (2s and 2p) has a **total of *8 electrons***. And, the third (last) shell has ***1 electron***.

The shell diagram of the Na atom is shown below. The shell nearest the nucleus (first shell) has 2 electrons (2 dots), the second shell has 8 electrons and the last (outermost) shell has 1 electron. (**2.8.1**)



Concept Review Exercises

1. How are electrons organized in atoms?
2. What information does an electron configuration convey?
3. What is the difference between core electrons and valence electrons?

Answers

1. Electrons are organized into shells and subshells around nuclei.
2. The electron configuration states the arrangement of electrons in shells and subshells.
3. Valence electrons are in the highest-numbered shell; all other electrons are core electrons.

Key Takeaway

- Electrons are organized into shells and subshells about the nucleus of an atom.
- The valence electrons determine the reactivity of an atom.

Exercises

4.10: Looking for Patterns - The Periodic Table

Learning Objectives

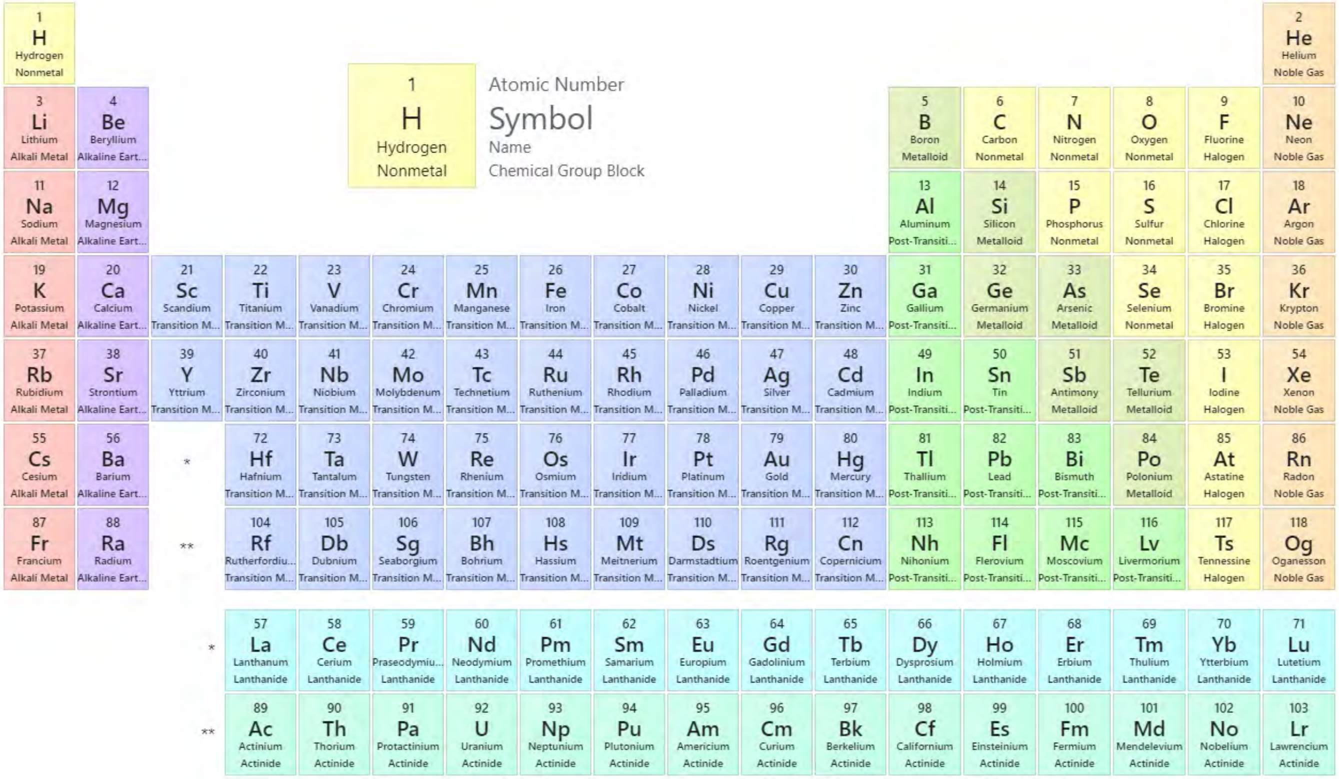
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

> *Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.*

The periodic table is one of the cornerstones of chemistry because it organizes all of the known elements on the basis of their chemical properties. A modern version is shown in Figure 4.10.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.



1 H Hydrogen Nonmetal
Atomic Number
Symbol
Name
Chemical Group Block

1 H Hydrogen Nonmetal																	2 He Helium Noble Gas
3 Li Lithium Alkali Metal	4 Be Beryllium Alkaline Eart...											5 B Boron Metalloid	6 C Carbon Nonmetal	7 N Nitrogen Nonmetal	8 O Oxygen Nonmetal	9 F Fluorine Halogen	10 Ne Neon Noble Gas
11 Na Sodium Alkali Metal	12 Mg Magnesium Alkaline Eart...											13 Al Aluminum Post-Transiti...	14 Si Silicon Metalloid	15 P Phosphorus Nonmetal	16 S Sulfur Nonmetal	17 Cl Chlorine Halogen	18 Ar Argon Noble Gas
19 K Potassium Alkali Metal	20 Ca Calcium Alkaline Eart...	21 Sc Scandium Transition M...	22 Ti Titanium Transition M...	23 V Vanadium Transition M...	24 Cr Chromium Transition M...	25 Mn Manganese Transition M...	26 Fe Iron Transition M...	27 Co Cobalt Transition M...	28 Ni Nickel Transition M...	29 Cu Copper Transition M...	30 Zn Zinc Transition M...	31 Ga Gallium Post-Transiti...	32 Ge Germanium Metalloid	33 As Arsenic Metalloid	34 Se Selenium Nonmetal	35 Br Bromine Halogen	36 Kr Krypton Noble Gas
37 Rb Rubidium Alkali Metal	38 Sr Strontium Alkaline Eart...	39 Y Yttrium Transition M...	40 Zr Zirconium Transition M...	41 Nb Niobium Transition M...	42 Mo Molybdenum Transition M...	43 Tc Technetium Transition M...	44 Ru Ruthenium Transition M...	45 Rh Rhodium Transition M...	46 Pd Palladium Transition M...	47 Ag Silver Transition M...	48 Cd Cadmium Transition M...	49 In Indium Post-Transiti...	50 Sn Tin Post-Transiti...	51 Sb Antimony Metalloid	52 Te Tellurium Metalloid	53 I Iodine Halogen	54 Xe Xenon Noble Gas
55 Cs Cesium Alkali Metal	56 Ba Barium Alkaline Eart...	*	72 Hf Hafnium Transition M...	73 Ta Tantalum Transition M...	74 W Tungsten Transition M...	75 Re Rhenium Transition M...	76 Os Osmium Transition M...	77 Ir Iridium Transition M...	78 Pt Platinum Transition M...	79 Au Gold Transition M...	80 Hg Mercury Transition M...	81 Tl Thallium Post-Transiti...	82 Pb Lead Post-Transiti...	83 Bi Bismuth Post-Transiti...	84 Po Polonium Metalloid	85 At Astatine Halogen	86 Rn Radon Noble Gas
87 Fr Francium Alkali Metal	88 Ra Radium Alkaline Eart...	**	104 Rf Rutherfordiu... Transition M...	105 Db Dubnium Transition M...	106 Sg Seaborgium Transition M...	107 Bh Bohrium Transition M...	108 Hs Hassium Transition M...	109 Mt Meitnerium Transition M...	110 Ds Darmstadtium Transition M...	111 Rg Roentgenium Transition M...	112 Cn Copernicium Transition M...	113 Nh Nihonium Post-Transiti...	114 Fl Flerovium Post-Transiti...	115 Mc Moscovium Post-Transiti...	116 Lv Livermorium Post-Transiti...	117 Ts Tennessine Halogen	118 Og Oganesson Noble Gas
		*	57 La Lanthanum Lanthanide	58 Ce Cerium Lanthanide	59 Pr Praseodymiu... Lanthanide	60 Nd Neodymium Lanthanide	61 Pm Promethium Lanthanide	62 Sm Samarium Lanthanide	63 Eu Europium Lanthanide	64 Gd Gadolinium Lanthanide	65 Tb Terbium Lanthanide	66 Dy Dysprosium Lanthanide	67 Ho Holmium Lanthanide	68 Er Erbium Lanthanide	69 Tm Thulium Lanthanide	70 Yb Ytterbium Lanthanide	71 Lu Lutetium Lanthanide
		**	89 Ac Actinium Actinide	90 Th Thorium Actinide	91 Pa Protactinium Actinide	92 U Uranium Actinide	93 Np Neptunium Actinide	94 Pu Plutonium Actinide	95 Am Americium Actinide	96 Cm Curium Actinide	97 Bk Berkelium Actinide	98 Cf Californium Actinide	99 Es Einsteinium Actinide	100 Fm Fermium Actinide	101 Md Mendelevium Actinide	102 No Nobelium Actinide	103 Lr Lawrencium Actinide

Figure 4.10.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

> *The word halogen comes from the Greek for "salt maker" because these elements combine with other elements to form a group of compounds called salts.*

To Your Health: Radon

Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building's inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

Each row of elements on the periodic table is called a period. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 4.10.2. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 4.10.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

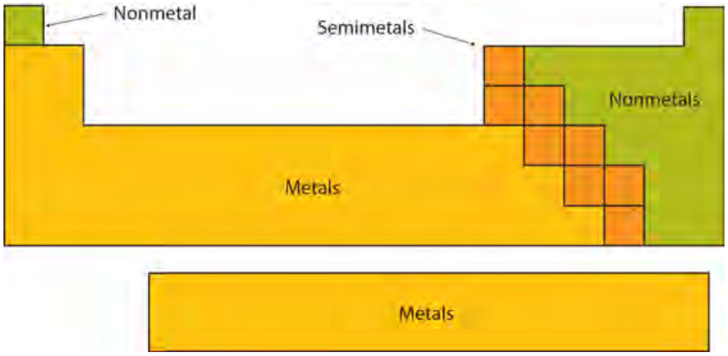


Figure 4.10.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

✓ Example 4.10.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

a. Se

b. Mg
c. Ge

Solution

a. In Figure 4.10.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

? Exercise 4.10.1

Based on its location in the periodic table, do you expect indium to be a nonmetal, a metal, or a metalloid?

Answer

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure 4.10.3. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

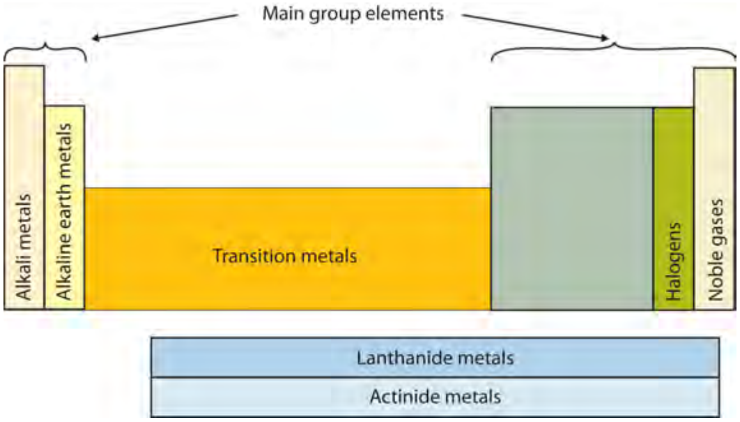


Figure 4.10.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth’s crust, respectively.

Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

Group 2: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for “salt forming,” which reflects that all of the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth’s ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

✓ Example 4.10.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

a. Lithium is an alkali metal (Group 1)
b. Argon is a noble gas (Group 18)
c. Chlorine is a halogen (Group 17)

? Exercise 4.10.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a:

Fluorine is a halogen (Group 17).

Answer b:

Calcium is a alkaline earth metal (Group 2).

Answer c:

Krypton is a noble gas (Group 18).

✓ Example 4.10.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. Li
b. Ar
c. Am
d. Fe

Solution

a. Lithium is a metal.
b. Argon is a non metal.
c. Americium is an inner transition metal.
d. Iron is a transition metal.

? Exercise 4.10.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. F
b. U
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

Uranium is a metal (and a inner transition metal too).

Answer c:

Copper is a metal (and a transition metal too).

Summary

The periodic table is an arrangement of the elements in order of increasing atomic number. Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods. Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.

The elements can be broadly divided into metals, nonmetals, and semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

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4.S: Elements, Atoms, and the Periodic Table (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms and ask yourself how they relate to the topics in the chapter.

An element is a substance that cannot be broken down into simpler chemical substances. Only about 90 naturally occurring elements are known. They have varying abundances on Earth and in the body. Each element has a one- or two-letter **chemical symbol**.

The **modern atomic theory** states that the smallest piece of an element is an **atom**. Individual atoms are extremely small, on the order of 10^{-10} m across. Most elements exist in pure form as individual atoms, but some exist as **diatomic molecules**. Atoms themselves are composed of subatomic particles. The **electron** is a tiny subatomic particle with a negative charge. The **proton** has a positive charge and, while small, is much larger than the electron. The **neutron** is also much larger than an electron but has no electrical charge.

Protons, neutrons, and electrons have a specific arrangement in an atom. The protons and neutrons are found in the center of the atom, grouped together into a **nucleus**. The electrons are found in fuzzy clouds around the nucleus.

Each element has a characteristic number of protons in its nucleus. This number of protons is the **atomic number** of the element. An element may have different numbers of neutrons in the nuclei of its atoms; such atoms are referred to as **isotopes**. Two isotopes of hydrogen are deuterium, with a proton and a neutron in its nucleus, and tritium, with a proton and two neutrons in its nucleus. The sum of the numbers of protons and neutrons in a nucleus is called the **mass number** and is used to distinguish isotopes from each other.

Masses of individual atoms are measured in **atomic mass units**. An atomic mass unit is equal to 1/12th of the mass of a single carbon-12 atom. Because different isotopes of an element have different masses, the **atomic mass** of an element is a weighted average of the masses of all the element's naturally occurring isotopes.

The modern theory of electron behavior is called **quantum mechanics**. According to this theory, electrons in atoms can only have specific, or **quantized**, energies. Electrons are grouped into general regions called **shells**, and within these into more specific regions called **subshells**. There are four types of subshells, and each type can hold up to a maximum number of electrons. The distribution of electrons into shells and subshells is the **electron configuration** of an atom. Chemistry typically occurs because of interactions between the electrons of the outermost shell of different atoms, called the valence shell electrons. Electrons in inner shells are called core electrons.

Elements are grouped together by similar chemical properties into a chart called the **periodic table**. Vertical columns of elements are called **groups** or **families**. Some of the groups of elements have names, like the alkali metals, the alkaline earth metals, the halogens, and the noble gases. A horizontal row of elements is called a **period**. Periods and groups have differing numbers of elements in them. The periodic table separates elements into **metals**, **nonmetals**, and **semimetals**. The periodic table is also separated into **main group elements**, **transition metals**, **lanthanide elements**, and **actinide elements**. The lanthanide and actinide elements are also referred to as **inner transition metal elements**. The shape of the periodic table reflects the sequential filling of shells and subshells in atoms.

The periodic table helps us understand trends in some of the properties of atoms. One such property is the **atomic radius** of atoms. From top to bottom of the periodic table, atoms get bigger because electrons are occupying larger and bigger shells. From left to right across the periodic table, electrons are filling the same shell but are being attracted by an increasing positive charge from the nucleus, and thus the atoms get smaller.

CHAPTER OVERVIEW

5: Nuclear Chemistry

Most chemists pay little attention to the nucleus of an atom except to consider the number of protons it contains because that determines an element’s identity. However, in nuclear chemistry, the composition of the nucleus and the changes that occur there are very important. Applications of nuclear chemistry may be more widespread than you realize. Many people are aware of nuclear power plants and nuclear bombs, but nuclear chemistry also has applications ranging from smoke detectors to medicine, from the sterilization of food to the analysis of ancient artifacts. In this chapter, we will examine some of the basic concepts of nuclear chemistry and some of the nuclear reactions that are important in our everyday lives.

5.1: Prelude to Nuclear Chemistry

Most of us may not be aware of a device in our homes that guards our safety and, at the same time, depends on radioactivity to operate properly. This device is a smoke detector.

A typical smoke detector contains an electric circuit that includes two metal plates about 1 cm apart. A battery in the circuit creates a voltage between the plates. Next to the plates is a small disk containing a tiny amount (~0.0002 g) of the radioactive element americium (Am). The radioactivity of the americium ionizes the air between the plates, causing a tiny current to constantly flow between them. (This constant drain on the battery explains why the batteries in smoke detectors should be replaced on a regular basis, whether the alarm has been triggered or not.)

Figure 5.1.1: Many people think of nuclear chemistry in connection with the nuclear power industry and atomic bombs but do not realize that most smoke detectors rely on nuclear chemistry and save countless lives every year. The applications of nuclear chemistry may be more widespread than you think. (CC BY-SA 3.0; Wile e2005).

When particles of smoke from a fire enter the smoke detector, they interfere with the ions between the metal plates, interrupting the tiny flow of current. When the current drops beneath a set value, another circuit triggers a loud alarm, warning of the possible presence of fire.

Although radioactive, the americium in a smoke detector is embedded in plastic and is not harmful unless the plastic package is taken apart, which is unlikely. Although many people experience an unfounded fear of radioactivity, smoke detectors provide an application of radioactivity that saves thousands of lives every year.

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5.2: Radioactivity

Learning Objectives

- To define and give examples of the major types of radioactivity.

Atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of the atom, while electrons circle the nucleus in shells and subshells and account for an atom's size. Remember, the notation for succinctly representing an isotope of a particular atom:

$$^{12}_{6}\mathrm{C} \tag{5.2.1}$$

The element in this example, represented by the symbol *C*, is carbon. Its atomic number, 6, is the lower left subscript on the symbol and is the number of protons in the atom. The mass number, the superscript to the upper left of the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 12, which means that the number of neutrons in the atom is 12 − 6 = 6 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neutrons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. The two isotopes of hydrogen, ^{2}H and ^{3}H, are given their own names: deuterium (D) and tritium (T), respectively. Another way of expressing a particular isotope is to list the mass number after the element name, like carbon-12 or hydrogen-3.

Atomic theory in the 19th century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source as the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an alpha particle, which is symbolized by the Greek letter α. An alpha particle is composed of two protons and two neutrons, and so it is the same as a helium nucleus. (We often use $^{4}_{2}\mathrm{He}$ to represent an alpha particle.) It has a 2+ charge. When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by two (because of the loss of two protons), and its mass number decreases by four (because of the loss of four nuclear particles). We can represent the emission of an alpha particle with a chemical equation—for example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{231}_{90}\mathrm{Th} \tag{5.2.2}$$

How do we know that a product of the reaction is $^{231}_{90}\mathrm{Th}$? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means we must have the same number of protons and neutrons on both sides of the chemical equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose 4 nuclear particles of the original 235, there are 231 remaining. Thus, we use subtraction to identify the isotope of the thorium atom—in this case, $^{231}_{90}\mathrm{Th}$.

Chemists often use the names *parent isotope* and *daughter isotope* to represent the original atom and the product other than the alpha particle. In the previous example, $^{235}_{92}\mathrm{U}$ is the parent isotope, and $^{231}_{90}\mathrm{Th}$ is the daughter isotope. When one element changes into another in this manner, it undergoes *radioactive decay*.

Example 5.2.1

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as $^{222}_{86}\mathrm{Rn}$. We represent the alpha particle as $^{4}_{2}\mathrm{He}$ and use subtraction (222 − 4 = 218 and 86 − 2 = 84) to identify the daughter isotope as an isotope of polonium, $^{218}_{84}\mathrm{Po}$:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{218}_{84}\mathrm{Po}$$

? Exercise ? 5.2.1

Write the nuclear equation that represents the radioactive decay of polonium-209 by alpha particle emission and identify the daughter isotope.

Answer

$$^{209}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{205}_{82}\text{Pb}$$

The second major type of radioactive emission is called a beta particle, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1− charge. We can also represent a beta particle as ${}^{0}_{-1}\text{e}$ or β^-. The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e} \tag{5.2.3}$$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the electron is assigned an “atomic number” of 1−, equal to its charge.)

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called gamma rays, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*). For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He} + \gamma \tag{5.2.4}$$

This may not seem like much energy, but if 1 mol of radon atoms were to decay, the gamma ray energy would be 49 million kJ!

✓ Example 5.2.2

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is ${}^{12}_{5}\text{B}$ while one of the products is an electron, ${}^{0}_{-1}\text{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus, the complete nuclear equation is as follows:

$$^{12}_{5}\text{B} \rightarrow {}^{12}_{6}\text{C} + {}^{0}_{-1}\text{e} + \gamma$$

The daughter isotope is ${}^{12}_{6}\text{C}$.

? Exercise 5.2.2

Write the nuclear equation that represents the radioactive decay of iodine-131 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Answer

$$^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}\text{e} + \gamma$$

Alpha, beta, and gamma emissions have different abilities to penetrate matter. The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter. Table 5.2.1 summarizes the properties of the three main types of radioactive emissions.

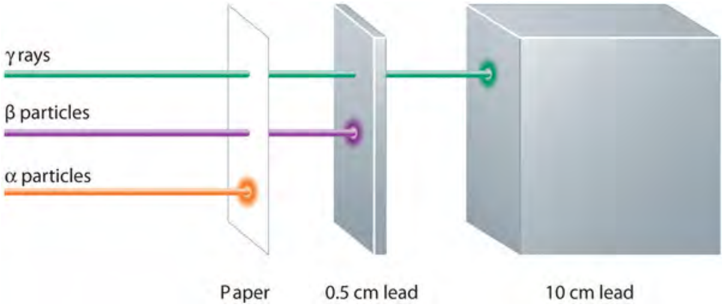

Figure 5.2.2: Different emissions exhibit different pentration powers. (CC BY-NC-SA 3.0; anonymous)

Table 5.2.1: The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, $^4_2\mathrm{He}$	β, $^{0}_{-1}\mathrm{e}$	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1−	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called *spontaneous fission* (or fission). Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\mathrm{U} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 2\,{}^{1}_{0}\mathrm{n} \tag{5.2.5}$$

where $^1_0\mathrm{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and the mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208.

> *Fission is the radioactive process used in nuclear power plants and one type of nuclear bomb.*

Key Takeaway

The major types of radioactivity include alpha particles, beta particles, and gamma rays.

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5.3: Half-Life

Learning Objectives

- To define *half-life*.
- To determine the amount of radioactive substance remaining after a given number of half-lives.

Whether or not a given isotope is radioactive is a characteristic of that particular isotope. Some isotopes are stable indefinitely, while others are radioactive and decay through a characteristic form of emission. As time passes, less and less of the radioactive isotope will be present, and the level of radioactivity decreases. An interesting and useful aspect of radioactive decay is the **half-life**. The half-life of a radioactive isotope is the amount of time it takes for one-half of the radioactive isotope to decay. The half-life of a specific radioactive isotope is constant; it is unaffected by conditions and is independent of the initial amount of that isotope.

Consider the following example. Suppose we have 100.0 g of ^3H (tritium, a radioactive isotope of hydrogen). It has a half-life of 12.3 y. After 12.3 y, half of the sample will have decayed to ^3He by emitting a beta particle, so that only 50.0 g of the original ^3H remains. After another 12.3 y—making a total of 24.6 y—another half of the remaining ^3H will have decayed, leaving 25.0 g of ^3H. After another 12.3 y—now a total of 36.9 y—another half of the remaining ^3H will have decayed, leaving 12.5 g of ^3H. This sequence of events is illustrated in Figure 5.3.1.

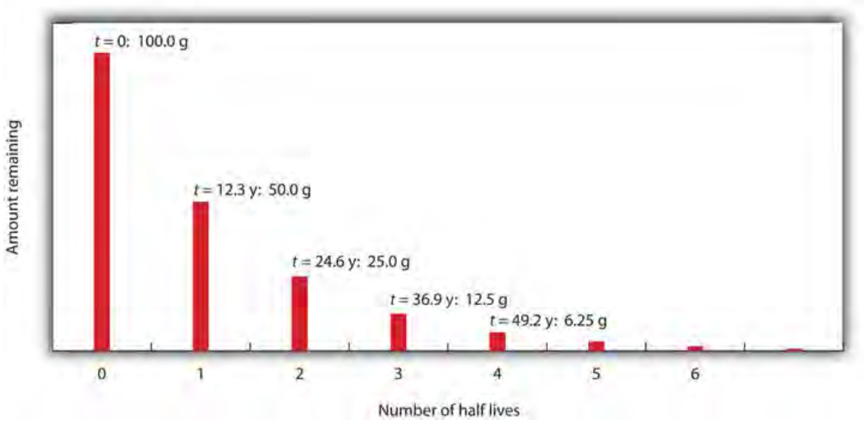


Figure 5.3.1: Radioactive Decay. During each successive half-life, half of the initial amount will radioactively decay.

We can determine the amount of a radioactive isotope remaining after a given number half-lives by using the following expression:

$$\text{amount remaining} = \text{initial amount} \times \left(\frac{1}{2}\right)^{n} \tag{5.3.1}$$

where *n* is the number of half-lives. This expression works even if the number of half-lives is not a whole number.

Example 5.3.1

The half-life of ^{20}F is 11.0 s. If a sample initially contains 5.00 g of ^{20}F, how much ^{20}F remains after 44.0 s?

Solution

If we compare the time that has passed to the isotope's half-life, we note that 44.0 s is exactly 4 half-lives, so we can use Equation ??? with $n = 4$. Substituting and solving results in the following:

$$\begin{aligned} \text{amount remaining} &= 5.00\ g \times \left(\frac{1}{2}\right)^{4} \\ &= 5.00\ g \times \frac{1}{16} \\ &= 0.313\ g \end{aligned}$$

Less than one-third of a gram of ^{20}F remains.

? Exercise 5.3.2

The half-life of ^{44}Ti is 60.0 y. A sample initially contains 0.600 g of ^{44}Ti. How much ^{44}Ti remains after 180.0 y?

Answer

0.075 g.

Half-lives of isotopes range from fractions of a microsecond to billions of years. Table 5.3.1 lists the half-lives of some isotopes.

Table 5.3.1: Half-Lives of Various Isotopes

Isotope	Half-Life
^{3}H	12.3 y
^{14}C	5,730 y
^{40}K	1.26 × 10^{9} y
^{51}Cr	27.70 d
^{90}Sr	29.1 y
^{131}I	8.04 d
^{222}Rn	3.823 d
^{235}U	7.04 × 10^{8} y
^{238}U	4.47 × 10^{9} y
^{241}Am	432.7 y
^{248}Bk	23.7 h
^{260}Sg	4 ms

✓ Example 5.3.2

The isotope I-125 is used in certain laboratory procedures and has a half-life of 59.4 days. If the initial activity of a sample of I-125 is 32,000 counts per minute (cpm), how much activity will be present in 178.2 days?

Solution

We begin by determining how many half-lives are represented by 178.2 days:

$$\frac{178.2 \text{ days}}{59.4 \text{ days/half-life}} = 3 \text{ half-lives}$$

Then we simply count activity:

$$\begin{aligned} \text{initial activity } (t_0) &= 32,000 \text{ cpm} \\ \text{after one half-life } &= 16,000 \text{ cpm} \\ \text{after two half-lives } &= 8,000 \text{ cpm} \\ \text{after three half-lives} &= 4,000 \text{ cpm} \end{aligned}$$

Be sure to keep in mind that the initial count is at time zero (t_0) and we subtract from that count at the first half-life. The second half-life has an activity of half the previous count (not the initial count).

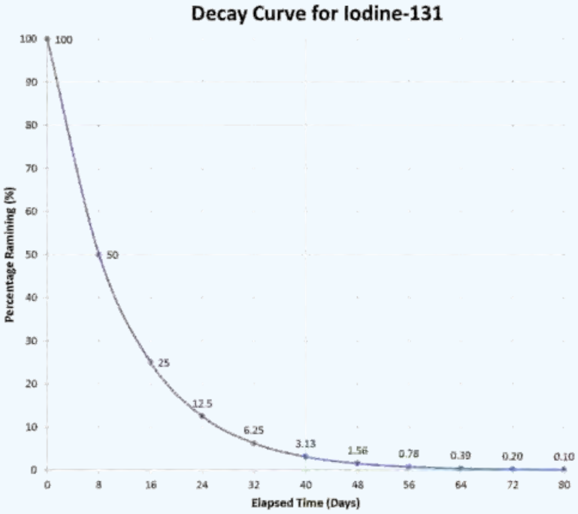
Equation 5.3.1 can be used to calculate the amount of radioactivity remaining after a given time:

$$N_t = N_0 \times (0.5)^{\text{number of half-lives}}$$

where $N_t =$ activity at time t and $N_0 =$ initial activity at time $t = 0$.

If we have an initial activity of $42,000$ cpm, what will the activity be after four half-lives?

$$\begin{aligned} N_t &= N_0(0.5)^4 \\ &= (42,000)\,(0.5)\,(0.5)\,(0.5)\,(0.5) \\ &= 2625 \text{ cpm} \end{aligned}$$



Typical radioactive decay curve.

The graph above illustrates a typical decay curve for $\text{I}-125$. The activity decreases by one-half during each succeeding half-life.

rcise

A sample of Ac-225 originally contained 80 grams and after 50 days only 2.5 grams of the original Ac-225 remain. What is the half life of Ac-225?

Answer

10 days

Looking Closer: Half-Lives of Radioactive Elements

Many people think that the half-life of a radioactive element represents the amount of time an element is radioactive. In fact, it is the time required for half—not all—of the element to decay radioactively. Occasionally, however, the daughter element is also radioactive, so its radioactivity must also be considered.

The expected working life of an ionization-type smoke detector (described in the opening essay) is about 10 years. In that time, americium-241, which has a half-life of about 432 y, loses less than 4% of its radioactivity. A half-life of 432 y may seem long to us, but it is not very long as half-lives go. Uranium-238, the most common isotope of uranium, has a half-life of about 4.5×10^9 y, while thorium-232 has a half-life of 14×10^9 y.

On the other hand, some nuclei have extremely short half-lives, presenting challenges to the scientists who study them. The longest-lived isotope of lawrencium, ^{262}Lr, has a half-life of 3.6 h, while the shortest-lived isotope of lawrencium, ^{252}Lr, has a half-life of 0.36 s. As of this writing, the largest atom ever detected has atomic number 118, mass number 293, and a half-life of 120 ns. Can you imagine how quickly an experiment must be done to determine the properties of elements that exist for so short a time?

Key Takeaways

- Natural radioactive processes are characterized by a half-life, the time it takes for half of the material to decay radioactively.
- The amount of material left over after a certain number of half-lives can be easily calculated.

5.4: Units of Radioactivity

Learning Objectives

- To express amounts of radioactivity in a variety of units.

Previously, we used mass to indicate the amount of radioactive substance present. This is only one of several units used to express amounts of radiation. Some units describe the number of radioactive events occurring per unit time, while others express the amount of a person's exposure to radiation.

Perhaps the direct way of reporting radioactivity is the number of radioactive decays per second. One decay per second is called one becquerel (Bq). Even in a small mass of radioactive material, however, there are many thousands of decays or disintegrations per second. The unit curie (Ci), now defined as 3.7×10^{10} decays per second, was originally defined as the number of decays per second in 1 g of radium. Many radioactive samples have activities that are on the order of microcuries (µCi) or more. Both the becquerel and curie can be used in place of grams to describe quantities of radioactive material. As an example, the amount of americium in an average smoke detector has an activity of 0.9 µCi.

> *The unit becquerel is named after Henri Becquerel, who discovered radioactivity in 1896. The unit curie is named after Polish scientist Marie Curie, who performed some of the initial investigations into radioactive phenomena and discovered the elements, polonium (Po) and radium (Ra) in the early 1900s.*

Example 5.4.1

A sample of radium has an activity of 16.0 mCi (millicuries). If the half-life of radium is 1,600 y, how long before the sample's activity is 1.0 mCi?

Solution

The following table shows the activity of the radium sample over multiple half-lives:

Solutions to Example 11.3.1

Time in Years	Activity
0	16.0 mCi
1,600	8.0 mCi
3,200	4.0 mCi
4,800	2.0 mCi
6,400	1.0 mCi

Over a period of 4 half-lives, the activity of the radium will be halved four times, at which point its activity will be 1.0 mCi. Thus, it takes 4 half-lives, or 4 × 1,600 y = 6,400 y, for the activity to decrease to 1.0 mCi.

Exercise 5.4.1

A sample of radon has an activity of 60,000 Bq. If the half-life of radon is 15 h, how long before the sample's activity is 3,750 Bq?

Answer

60 hrs.

Other measures of radioactivity are based on the effects it has on living tissue. Radioactivity can transfer energy to tissues in two ways: through the kinetic energy of the particles hitting the tissue and through the electromagnetic energy of the gamma rays being absorbed by the tissue. Either way, the transferred energy—like thermal energy from boiling water—can damage the tissue.

The rad (an acronym for radiation absorbed dose) is a unit equivalent to a gram of tissue absorbing 0.01 J:

$$1 \text{ rad} = 0.01 \text{ J/g}$$

Another unit of radiation absorption is the gray (Gy):

$$1 \text{ Gy} = 100 \text{ rad}$$

The rad is more common. To get an idea of the amount of energy this represents, consider that the absorption of 1 rad by 70,000 g of H_2O (approximately the same mass as a 150 lb person) would increase its temperature by only 0.002°C. This may not seem like a lot, but it is enough energy to break about 1×10^{21} molecular C–C bonds in a person's body. That amount of damage would not be desirable.

Predicting the effects of radiation is complicated by the fact that various tissues are affected differently by different types of emissions. To quantify these effects, the unit **rem** (an acronym for roentgen equivalent, man) is defined as

$$\text{rem} = \text{rad} \times \text{RBE}$$

where RBE is the *relative biological effectiveness factor* is a number greater than or equal to 1 that takes into account the type of radioactive emission and sometimes the type of tissue being exposed. For beta particles, RBE factor equals 1. For alpha particles striking most tissues, the factor is 10, but for eye tissue, the factor is 30. Most radioactive emissions that people are exposed to are on the order of a few dozen millirems (mrem) or less; a medical X ray is about 20 mrem. A sievert (Sv) is a related unit and is defined as 100 rem.

What is a person's annual exposure to radioactivity and radiation? Table 5.4.1 lists the sources and annual amounts of radiation exposure. It may surprise you to learn that fully 82% of the radioactivity and radiation exposure we receive is from natural sources—sources we cannot avoid. Fully 10% of the exposure comes from our own bodies—largely from ^{14}C and ^{40}K.

Table 5.4.1: Average Annual Radiation Exposure (Approximate)

Source	**Amount (mrem)**
radon gas	200
medical sources	53
radioactive atoms in the body naturally	39
terrestrial sources	28
cosmic sources	28
consumer products	10
nuclear energy	0.05
Total	358

> *Flying from New York City to San Francisco adds 5 mrem to your overall radiation exposure because the plane flies above much of the atmosphere, which protects us from most cosmic radiation.*

The actual effects of radioactivity and radiation exposure on a person's health depend on the type of radioactivity, the length of exposure, and the tissues exposed. Table 5.4.2 lists the potential threats to health at various amounts of exposure over short periods of time (hours or days).

Table 5.4.2: Effects of Short-Term Exposure to Radioactivity and Radiation

Exposure (rem)	Effect
1 (over a full year)	no detectable effect
∼20	increased risk of some cancers

Exposure (rem)	Effect
∼100	damage to bone marrow and other tissues; possible internal bleeding; decrease in white blood cell count
200–300	visible “burns” on skin, nausea, vomiting, and fatigue
>300	loss of white blood cells; hair loss
∼600	death

One of the simplest ways of detecting radioactivity is by using a piece of photographic film embedded in a badge or a pen. On a regular basis, the film is developed and checked for exposure. A comparison of the exposure level of the film with a set of standard exposures indicates the amount of radiation a person was exposed to.

Figure 5.4.1: Detecting Radioactivity. A Geiger counter is a common instrument used to detect radioactivity.

Another means of detecting radioactivity is an electrical device called a Geiger counter (Figure 5.4.1). It contains a gas-filled chamber with a thin membrane on one end that allows radiation emitted from radioactive nuclei to enter the chamber and knock electrons off atoms of gas (usually argon). The presence of electrons and positively charged ions causes a small current, which is detected by the Geiger counter and converted to a signal on a meter or, commonly, an audio circuit to produce an audible “click.”

Key Takeaway

- Radioactivity can be expressed in a variety of units, including rems, rads, and curies.

5.5: Uses of Radioactive Isotopes

Learning Objectives

- To learn some applications of radioactivity.

Radioactive isotopes have a variety of applications. Generally, however, they are useful either because we can detect their radioactivity or we can use the energy they release.

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A tracer is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some tritium-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium, ^{3}H, is a radioactive isotope of hydrogen.)

Tracers can also be used to follow the steps of a complex chemical reaction. After incorporating radioactive atoms into reactant molecules, scientists can track where the atoms go by following their radioactivity. One excellent example of this is the use of radioactive carbon-14 to determine the steps involved in the photosynthesis in plants. We know these steps because researchers followed the progress of the radioactive carbon-14 throughout the process.

Radioactive isotopes are useful for establishing the ages of various objects. The half-life of radioactive isotopes is unaffected by any environmental factors, so the isotope acts like an internal clock. For example, if a rock is analyzed and is found to contain a certain amount of uranium-235 and a certain amount of its daughter isotope, we can conclude that a certain fraction of the original uranium-235 has radioactively decayed. If half of the uranium has decayed, then the rock has an age of one half-life of uranium-235, or about 4.5×10^9 y. Many analyses like this, using a wide variety of isotopes, have indicated that the age of Earth itself is over 4×10^9 y. In another interesting example of radioactive dating, ^{3}H dating has been used to verify the stated vintages of some old fine wines.

Carbon-14 (half-life is 5,370 y) is particularly useful in determining the age of once-living artifacts (e.g., animal or plant matter). A tiny amount of carbon-14 is produced naturally in the upper reaches of the atmosphere, and living things incorporate some of it into their tissues, building up to a constant, although very low, level. Once a living thing dies, however, it no longer acquires carbon-14, and as time passes, the carbon-14 that was in the tissues decays. If a once-living artifact is discovered and analyzed many years after its death, with the remaining carbon-14 compared to the known constant level, an approximate age of the artifact can be determined. Using such methods, scientists determined that the age of the Shroud of Turin (made of linen, which comes from the flax plant, and purported by some to be the burial cloth of Jesus Christ; Figure 5.5.1) is about 600–700 y, not 2,000 y as claimed by some. Scientists were also able to use radiocarbon dating to show that the age of a mummified body found in the ice of the Alps was 5,300 y.

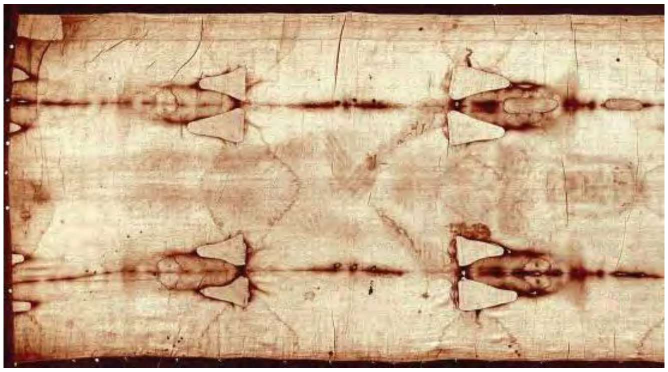


Figure 5.5.1: Shroud of Turin. In 1989, several groups of scientists used carbon-14 dating to demonstrate that the age of the Shroud of Turin was only 600–700 y. Many people still cling to a different notion, despite the scientific evidence.

The radiation emitted by some radioactive substances can be used to kill microorganisms on a variety of foodstuffs, which extends the shelf life of these products. Produce such as tomatoes, mushrooms, sprouts, and berries are irradiated with the emissions from coba . This exposure kills a lot of the bacteria that cause spoilage, so the produce stays fresh longer. Eggs and some meat, such as beef, pork, and poultry, can also be irradiated. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.

Radioactive isotopes have numerous medical applications—diagnosing and treating illnesses and diseases. One example of a diagnostic application is using radioactive iodine-131 to test for thyroid activity (Figure $5.5.2$). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of iodine-131 is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

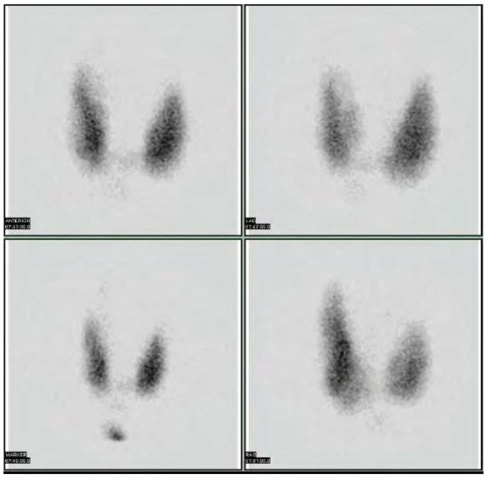


Figure $5.5.2$: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, Wikipedia, Thyroid Scan(opens in new window) [en.Wikipedia.org].

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect. However, therapeutic applications usually require much larger doses because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor is detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less then 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect but also ease the pain of bone cancers. Table $5.5.1$ lists several radioactive isotopes and their medical uses.

Table $5.5.1$: Some Radioactive Isotopes That Have Medical Applications

Isotope	Use
^{32}P	cancer detection and treatment, especially in eyes and skin
^{59}Fe	anemia diagnosis
^{60}Co	gamma ray irradiation of tumors
^{99m}Tc	brain, thyroid, liver, bone marrow, lung, heart, and intestinal scanning; blood volume determination
^{131}I	diagnosis and treatment of thyroid function
^{133}Xe	lung imaging
^{198}Au	liver disease diagnosis

In addition to the direct application of radioactive isotopes to diseased tissue, the gamma ray emissions of some isotopes can be directed toward the tissue to be destroyed. Cobalt-60 is a useful isotope for this kind of procedure.

To Your Health: Positron Emission Tomography Scans

One relatively rare form of radioactivity is called *positron emission*. It is similar to beta particle emission, except that instead of emitting an electron, a nucleus emits a positively charged electron, called a *positron*. A positron is actually a piece of antimatter; therefore, when a positron encounters an electron, both particles are converted into high-energy gamma radiation.

Isotopes that emit positrons can be employed in a medical imaging technique called *positron emission tomography* (*PET*). A patient receives a compound containing a positron-emitting isotope, either intravenously or by ingestion. The radioactive compound travels throughout the body, and the patient is then pushed slowly through a ring of sensors that detect the gamma radiation given off by the annihilation of positrons and electrons. A computer connected to the sensors constructs a three-dimensional image of the interior of part or all of the patient's body, allowing doctors to see organs or tumors or regulate the function of various organs (such as the brain or the heart) to diagnose the medical condition of the patient.

Figure 5.5.3: (left) Combined apparatus for positron emission tomography (PET) and X-ray computer tomography (CT), Siemens Biograph (right) Whole-body PET scan using 18F-FDG. Images used with permission from Wikipedia.

Two isotopes that undergo positron emission are carbon-11 and fluorine-18, with half-lives of 20.4 and 110 min, respectively. Both isotopes can be incorporated into sugar molecules and introduced into the body. Doctors can use the intensity of gamma ray emission to find tissues that metabolize the sugar faster than other tissues; fast-metabolizing tissue is one sign of a malignant (i.e., cancerous) tumor. Researchers use similar techniques to map areas of the brain that are most active during specific tasks, such as reading or speaking.

PET is one of many diagnostic and treatment methods that physicians use to improve the quality of our lives. It is one of the many positive uses of radioactivity in society.

Key Takeaway

- Radioactivity has several practical applications, including tracers, medical applications, dating once-living objects, and the preservation of food.

5.6: Nuclear Energy

Learning Objectives

- Explain where nuclear energy comes from.
- Describe the difference between fission and fusion.

Nuclear changes occur with a simultaneous release of energy. Where does this energy come from? If we could precisely measure the masses of the reactants and the products of a nuclear reaction, we would notice that the amount of mass drops slightly in the conversion from reactants to products. Consider the following nuclear reaction, in which the molar mass of each species is indicated to four decimal places:

$$\underbrace{^{235}\text{U}}_{235.0439} \rightarrow \underbrace{^{139}\text{Ba}}_{138.9088} + \underbrace{^{94}\text{Kr}}_{93.9343} + \underbrace{2\,^{1}\text{n}}_{2\times 1.0087}$$

If we compare the mass of the reactant (235.0439) to the masses of the products (sum = 234.8605), we notice a mass difference of −0.1834 g, or −0.0001834 kg. Where did this mass go?

According to Albert Einstein’s theory of relativity, energy (E) and mass (m) are related by the following equation:

$$E = mc^2 \tag{5.6.1}$$

where c is the speed of light, or 3.00×10^8 m/s. In the course of the uranium nuclear chemical reaction, the mass difference is converted to energy, which is given off by the reaction:

$$\begin{aligned} E &= (-0.0001834\ kg)(3.00 \times 10^8\ m/s)^2 \\ &= -1.65 \times 10^{13}\ J \\ &= -1.65 \times 10^{10}\ kJ \end{aligned}$$

That is, 16.5 billion kJ of energy are given off every time 1 mol of uranium-235 undergoes this nuclear reaction. This is an extraordinary amount of energy. Compare it to combustion reactions of hydrocarbons, which give off about 650 kJ/mol of energy for every CH_2 unit in the hydrocarbon—on the order of *hundreds* of kilojoules per mole. Nuclear reactions give off *billions* of kilojoules per mole.

If this energy could be properly harvested, it would be a significant source of energy for our society. Nuclear energy involves the controlled harvesting of energy from fission reactions. The reaction can be controlled because the fission of uranium-235 (and a few other isotopes, such as plutonium-239) can be artificially initiated by injecting a neutron into a uranium nucleus. The overall nuclear equation, with energy included as a product, is then as follows:

$$^{235}\text{U} + {}^{1}\text{n} \longrightarrow {}^{139}\text{Ba} + {}^{94}\text{Kr} + 3\,{}^{1}\text{n} + \text{energy}$$

Thus, by the careful addition of extra neutrons into a sample of uranium, we can control the fission process and obtain energy that can be used for other purposes.

The Curie Family

Artificial or induced radioactivity was first demonstrated in 1934 by Irène Joliot-Curie and Frédéric Joliot, the daughter and son-in-law of Marie Curie.

Example 5.6.1

Plutonium-239 can absorb a neutron and undergo a fission reaction to produce an atom of gold-204 and an atom of phosphorus-31. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Solution

Using the data given, we can write the following initial equation:

$$_0^1\text{n} + {}_{94}^{239}\text{Pu} \rightarrow {}_{79}^{204}\text{Au} + {}_{15}^{31}\text{P} + ?_0^1\text{n}$$

In balanced nuclear equations, the sums of the subscripts on each sides of the equation are the same, as are the sums of the superscripts. The subscripts are already balanced: 0 + 94 = 94 and 79 + 15 = 94. The superscripts on the left equal 240 (1 + 239) but equal 235 (204 + 31) on the right. We need five more mass number units on the right. Five neutrons should be the products of the process for the mass numbers to balance. (Because the atomic number of a neutron is zero, including five neutrons on the right does not change the overall sum of the subscripts.) Thus, the balanced nuclear equation is as follows:

$$_0^1\text{n} + {}_{94}^{239}\text{Pu} \rightarrow {}_{79}^{204}\text{Au} + {}_{15}^{31}\text{P} + 5_0^1\text{n}$$

We predict that the overall process will give off five neutrons.

? Exercise 5.6.1

Uranium-238 can absorb a neutron and undergo a fission reaction to produce an atom of cesium-135 and an atom of rubidium-96. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Answer

$$_0^1\text{n} + {}_{92}^{238}\text{U} \rightarrow {}_{55}^{135}\text{Cs} + {}_{37}^{96}\text{Rb} + 8_0^1\text{n}$$

We predict that the overall process will give off eight neutrons.

A nuclear reactor is an apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes. Figure 5.6.1 shows a simplified diagram of a nuclear reactor. The energy from the controlled nuclear reaction converts liquid water into high-pressure steam, which is used to run turbines that generate electricity.

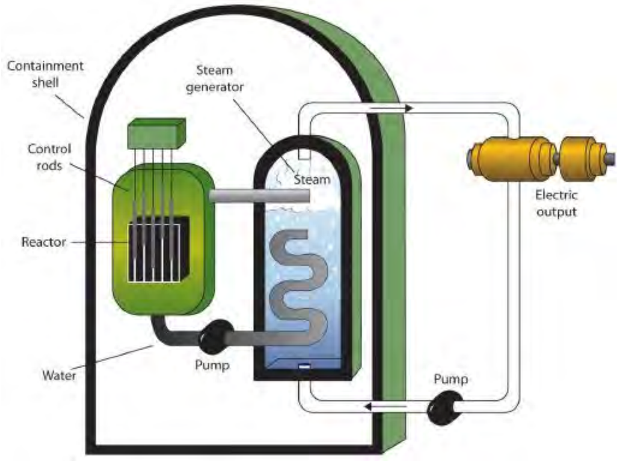


Figure 5.6.1: A Diagram of a Nuclear Power Plant for Generating Electricity. The two main components of the power plant are the nuclear reactor itself and the steam-driven turbine and electricity generator.

Notice that the fission of uranium produces two more free neutrons than were present to begin with. These neutrons can themselves stimulate other uranium nuclei to undergo fission, releasing yet more energy and even more neutrons, which can in turn induce even more uranium fission. A single neutron can thus begin a process that grows exponentially in a phenomenon called a chain reaction:

$$1 \rightarrow 2 \rightarrow 4 \rightarrow 8 \rightarrow 16 \rightarrow 32 \rightarrow 64 \rightarrow 128 \rightarrow 256 \rightarrow 512 \rightarrow 1{,}024 \rightarrow 2{,}048 \rightarrow 4{,}096 \rightarrow 8{,}192 \rightarrow 16{,}384 \rightarrow \ldots$$

Because energy is produced with each fission event, energy is also produced exponentially and in an uncontrolled fashion. The quick production of energy creates an explosion. This is the mechanism behind the atomic bomb.

> The first controlled chain reaction was achieved on December 2, 1942, in an experiment supervised by Enrico Fermi in a laboratory underneath the football stadium at the University of Chicago.

Although fairly simple in theory, an atomic bomb is difficult to produce, in part because uranium-235, the isotope that undergoes fission, makes up only 0.7% of natural uranium; the rest is mostly uranium-238, which does not undergo fission. (Remember that the radioactive process that a nucleus undergoes is characteristic of the isotope.) To make uranium useful for nuclear reactors, the uranium in uranium-235 must be *enriched* to about 3%. Enrichment of uranium is a laborious and costly series of physical and chemical separations. To be useful in an atomic bomb, the uranium in uranium-235 must be enriched to 70% or more. At lesser concentrations, the chain reaction cannot sustain itself, so no explosion is produced.

Fusion is another nuclear process that can be used to produce energy. In this process, smaller nuclei are combined to make larger nuclei, with an accompanying release of energy. One example is the hydrogen fusion, which makes helium. While the steps of the process are complicated, the net reaction is:

$$4\,{}^{1}\mathrm{H} \rightarrow {}^{4}\mathrm{He} + 2.58 \times 10^{12}\ J$$

Notice that the amount of energy given off per mole of reactant is only a fraction of the amount given off by the fission of 1 mol of uranium-235. On a mass (per gram) basis, however, the hydrogen fusion emits many times more energy than fission does. In addition, the product of fission is helium gas, not a wide range of isotopes (some of which are also radioactive) produced by fission.

The practical problem is that to perform fusion, extremely high pressures and temperatures are necessary. Currently, the only known stable systems undergoing fusion are the interiors of stars. The conditions necessary for fusion can be created using an atomic bomb, but the resulting fusion is uncontrollable (and the basis for another type of bomb, a hydrogen bomb). Currently, researchers are looking for safe, controlled ways of producing useful energy using fusion.

Career Focus: Nuclear Medicine Technologist

Generally speaking, a radiological technician deals with X ray equipment and procedures. A *nuclear medicine technologist* has similar responsibilities, using compounds containing radioactive isotopes to help diagnose and treat disease.

Nuclear medicine technologists administer the substances containing the radioactive isotope and subsequently operate the apparatus that detects the radiation produced by radioactive decay. The apparatus may be as simple as a piece of photographic film or as complex as a series of computer-controlled electronic detectors. The images obtained by the technologist are interpreted by a specially trained physician.

One of the chief responsibilities of a nuclear medicine technologist is safety. Improper exposure to radioactivity can be harmful to both patient and technologist alike. Therefore, the technologist must adhere to strict safety standards to keep unnecessary exposure as low as possible. The technologist must also know how to dispose of waste materials safely and appropriately.

Key Takeaways

- Nuclear energy comes from tiny mass changes in nuclei as radioactive processes occur.
- In fission, large nuclei break apart and release energy; in fusion, small nuclei merge together and release energy.

5.E: Nuclear Chemistry (Exercises)

11.1 Radioactivity

Concept Review Exercise

1. What are the major types of radioactivity? Write chemical equations demonstrating each type.

Answer

1. The major types of radioactivity are alpha decay, beta decay, and gamma ray emission; alpha decay with gamma emission: $^{222}_{86}\mathrm{Rn} \rightarrow\, ^{218}_{84}\mathrm{Po} +\, ^{4}_{2}\mathrm{He} + \gamma$; beta decay: $^{14}_{6}\mathrm{C} \rightarrow\, ^{14}_{7}\mathrm{N} +\, ^{0}_{-1}\mathrm{e}$ (answers will vary)

Exercises

1. Define *radioactivity*.
2. Give an example of a radioactive isotope.
3. How many protons and neutrons are in each isotope?
   a. $^{11}_{5}\mathrm{B}$
   b. $^{27}_{13}\mathrm{Al}$
   c. $^{56}\mathrm{Fe}$
   d. $^{224}\mathrm{Rn}$
4. How many protons and neutrons are in each isotope?
   a. $^{2}_{1}\mathrm{H}$
   b. $^{112}_{48}\mathrm{Cd}$
   c. $^{252}\mathrm{Es}$
   d. $^{40}\mathrm{K}$
5. Describe an alpha particle. What nucleus is it equivalent to?
6. Describe a beta particle. What subatomic particle is it equivalent to?
7. Explain what gamma rays are.
8. Explain why it is inappropriate to refer to gamma rays as gamma “particles.”
9. Plutonium has an atomic number of 94. Write the chemical equation for the alpha particle emission of $^{244}\mathrm{Pu}$. What is the daughter isotope?
10. Francium has an atomic number of 87. Write the chemical equation for the alpha particle emission of $^{212}\mathrm{Fr}$. What is the daughter isotope?
11. Tin has an atomic number of 50. Write the chemical equation for the beta particle emission of $^{121}\mathrm{Sn}$. What is the daughter isotope?
12. Technetium has an atomic number of 43. Write the chemical equation for the beta particle emission of $^{99}\mathrm{Tc}$. What is the daughter isotope?
13. Energies of gamma rays are typically expressed in units of megaelectron volts (MeV), where 1 MeV = 1.602×10^{-13} J. Using data provided in the text, calculate the energy, in megaelectron volts, of the gamma ray emitted when radon-222 decays.
14. The gamma ray emitted when oxygen-19 gives off a beta particle is 0.197 MeV. What is its energy in joules? (See Exercise 13 for the definition of a megaelectron volt.)
15. Which penetrates matter more deeply—alpha particles or beta particles? Suggest ways to protect yourself against both particles.
16. Which penetrates matter more deeply—alpha particles or gamma rays? Suggest ways to protect yourself against both emissions.
17. Define *nuclear fission*.
18. What general characteristic is typically necessary for a nucleus to undergo spontaneous fission?

Answers

1. Radioactivity is the spontaneous emission of particles and radiation from atomic nuclei.

2. C-14 or ^{14}C is an example of radioactive isotope (answers may vary).

3.
   a. 5 protons; 6 neutrons
   b. 13 protons; 14 neutrons
   c. 26 protons; 30 neutrons
   d. 86 protons; 138 neutrons

4.
   a. 1 proton; 1 neutron
   b. 48 protons; 64 neutrons
   c. 99 protons;153 neutrons
   d. 19 protons; 21 neutrons

5. An alpha particle is a combination of two protons and two neutrons and is equivalent to a helium nucleus.

6. A beta particle is an electron.

7. Gamma rays are high-energy electromagnetic radiation given off in radioactive decay.

8. Gamma rays have no mass. Hence not a particle.

9. ${}^{244}_{94}\mathrm{Pu} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{240}_{92}\mathrm{U}$; the daughter isotope is ${}^{240}_{92}\mathrm{U}$, an atom of uranium.

10. ${}^{212}_{87}\mathrm{Fr} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{208}_{85}\mathrm{At}$; the daughter isotope is ${}^{208}_{85}\mathrm{At}$, an atom of astatine.

11. ${}^{121}_{50}\mathrm{Sn} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{121}_{51}\mathrm{Sb}$; the daughter isotope is ${}^{121}_{51}\mathrm{Sb}$, an atom of antimony.

12. ${}^{99}_{43}\mathrm{Tc} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{99}_{44}\mathrm{Mo}$; the daughter isotope is ${}^{99}_{44}\mathrm{Mo}$, an atom of antimony.

13. 0.512 MeV

14. 3.16 x 10^{-14} J

15. Beta particles; shielding of the appropriate thickness can protect against both alpha and beta particles.

16. Gamma rays; can be shielded by thick, dense material such as lead (Pb). Alpha particles has low energy; can shielded by a piece of paper.

17. Nuclear fission is when large nuclei break down into smaller nuclei.

18. A nucleus must be very large. Examples are Th-232 and U-235.

11.2 Half-Life

Concept Review Exercises

1. Define *half-life*.
2. Describe a way to determine the amount of radioactive isotope remaining after a given number of half-lives.

Answers

1. Half-life is the amount of time needed for half of a radioactive material to decay.
2. take half of the initial amount for each half-life of time elapsed

Exercises

1. Do all isotopes have a half-life? Explain.

2. Which is more radioactive—an isotope with a long half-life or an isotope with a short half-life?

3. What percent of a sample remains after one half-life? Three half-lives?

4. The half-life of polonium-218 is 3.0 min. How much of a 0.540 mg sample would remain after 9.0 minutes have passed?

5. The half-life of protactinium-234 is 6.69 hours. If a 0.812 mg sample of Pa-239 decays for 40.14 hours, what mass of the isotope remains?

6. How long does it take for 1.00 g of ^{103}Pd to decay to 0.125 g if its half-life is 17.0 d?

7. How long does it take for 2.00 g of ^{94}Nb to decay to 0.0625 g if its half-life is 20,000 y?

8. It took 75 y for 10.0 g of a radioactive isotope to decay to 1.25 g. What is the half-life of this isotope?

9. It took 49.2 s for 3.000 g of a radioactive isotope to decay to 0.1875 g. What is the half-life of this isotope?

Answers

1. Only radioactive isotopes have half-lives.

2. An isotope with a shorter half-life decay more rapidly is more radioactive.

3. 1 half-life: 50%; 3 half-lives: 12.5%

4. 9.0 min = 3 half-lives (make 3 arrows): 0.540 mg --> 0.270 mg --> 0.135 mg --> 0.0675 mg

5. 0.0127 mg

6. 51.0 d

7. 100 000 y

8. 25 y

9. 12.3 s

11.3 Units of Radioactivity

Concept Review Exercise

1. What units are used to quantify radioactivity?

Answer

1. the curie, the becquerel, the rad, the gray, the sievert, and the rem

Exercises

1. Define *rad*.
2. Define *rem*.
3. How does a becquerel differ from a curie?
4. How is the curie defined?
5. A sample of radon gas has an activity of 140.0 mCi. If the half-life of radon is 1,500 y, how long before the activity of the sample is 8.75 mCi?
6. A sample of curium has an activity of 1,600 Bq. If the half-life of curium is 24.0 s, how long before its activity is 25.0 Bq?
7. If a radioactive sample has an activity of 65 µCi, how many disintegrations per second are occurring?

8. If a radioactive sample has an activity of 7.55 × 10^5 Bq, how many disintegrations per second are occurring?
9. Describe how a radiation exposure in rems is determined.
10. Which contributes more to the rems of exposure—alpha or beta particles? Why?
11. Use Table 11.3.2 to determine which sources of radiation exposure are inescapable and which can be avoided. What percentage of radiation is unavoidable?
12. What percentage of the approximate annual radiation exposure comes from radioactive atoms that are in the body naturally?
13. Explain how a film badge works to detect radiation.
14. Explain how a Geiger counter works to detect radiation.

Answers

1. Known as the radiation absorbed dose, a rad is the absorption of 0.01 J/g of tissue.

2. Known as roentgen equivalent man, a rem is an absorption of one rad times a factor. The factor is variable depending on the type of emission and the type of irradiated tissue.

3. A becquerel is smaller and equals 1 decay per second. A curie is 3.7 × 10^{10} Bq.

4. A curie is defined as 3.7 × 10^{10} decays per second.

5. 6000 y

6. 144 s

7. 2.41 × 10^6 disintegrations per second

8. 7.55 × 10^5 disintegrations per second

9. The radiation exposure is determined by the number of rads times the quality factor of the radiation.

10. Alpha contributes more than beta because of its bigger size and electrical charge.

11. At least 16% (terrestrial and cosmic sources) of radioactivity is unavoidable; the rest depends on what else a person is exposed to.

12. About 11% come from radioactive atoms that are in the body naturally.

13. A film badge uses film, which is exposed as it is subjected to radiation.

14. The Geiger counter consists of a tube with electrodes and is filled with an inert (argon) gas. Radiation entering the tube ionizes the gas, and the ions are attracted to the electrodes and produce an electric pulse (clicking sound).

11.4 Uses of Radioactive Isotopes

Concept Review Exercise

1. Describe some of the different ways that amounts of radioactivity are applied in society.

Answer

1. Radioactive isotopes are used in dating, as tracers, and in medicine as diagnostic and treatment tools.

Exercises

1. Define *tracer* is and give an example of how tracers work.
2. Name two isotopes that have been used as tracers.
3. Explain how radioactive dating works.
4. Name an isotope that has been used in radioactive dating.
5. The current disintegration rate for carbon-14 is 14.0 Bq. A sample of burnt wood discovered in an archaeological excavation is found to have a carbon-14 decay rate of 3.5 Bq. If the half-life of carbon-14 is 5,700 y, approximately how old is the wood sample?
6. A small asteroid crashes to Earth. After chemical analysis, it is found to contain 1 g of technetium-99 to every 3 g of ruthenium-99, its daughter isotope. If the half-life of technetium-99 is 210,000 y, approximately how old is the asteroid?
7. What do you think are some of the positive aspects of irradiation of food?
8. What do you think are some of the negative aspects of irradiation of food?
9. Describe how iodine-131 is used to both diagnose and treat thyroid problems.
10. List at least five organs that can be imaged using radioactive isotopes.
11. Which radioactive emissions can be used therapeutically?
12. Which isotope is used in therapeutics primarily for its gamma ray emissions?
13. What volume of a radioisotope should be given if a patient needs 125 mCi of a solution which contains 45 mCi in 5.0 mL?
14. Sodium-24 is used to treat leukemia. A 36-kg patient is prescribed 145 μCi/kg and it is supplied to the hospital in a vial containing 250 μCi/mL. What volume should be given to the patient?
15. Lead-212 is one of the radioisotopes used in the treatment of breast cancer. A patient needs a 15 μCi dose and it is supplied as a solution with a concentration of 2.5 μCi/mL. What volume does the patient need? Given the half-life of lead is 10.6 hours, what will be the radioactivity of the sample after approximately four days?

Answers

1. A tracer follows the path of a chemical or a physical process. One of the uses of a tracer is following the path of water underground (answers will vary).
2. Tritium (^{3}H) and Carbon-14 (^{14}C) (answers will vary)
3. Radioactive dating works by comparing the amounts of parent and daughter isotopes and calculating back to how long ago all of the material was just the parent isotope.
4. Carbon-14 (^{14}C) and Uranium-235 (^{235}U) (answers will vary)
5. about 11,400 y
6. about 420,000 y
7. increased shelf life (answers will vary)

8. reduction in the food's vitamin content and cost

9. Iodine-131 is preferentially absorbed by the thyroid gland and can be used to measure the gland's activity or destroy bad cells in the gland.

10. brain, bone, heart, thyroid, lung (answers will vary)

11. gamma rays, beta particles, or alpha particles

12. cobalt-60

13. 125mCi x (5.0mL/45mCi)=14mL

14. 36kg x (145μCi/kg) x (1mL/250μCi)=21mL

15. Volume given: 15μCi x (1mL/2.5μCi) = 6.0mL

Elapsed time in hours: 4 days x (24 hr/day) = 96 hr

Number of half-lives: 96 hrs/10.6 hours = 9

Radioactivity remaining after 9 half-lives: 0.029 μCi

11.5 Nuclear Energy

Concept Review Exercises

1. How is nuclear energy produced?
2. What is the difference between fission and fusion?

Answers

1. Nuclear energy is produced by carefully controlling the speed of a fission reaction.
2. In fission, large nuclei break down into small ones; in fusion, small nuclei combine to make larger ones. In both cases, a lot of energy is emitted.

Exercises

1. In the spontaneous fission of uranium-233, the following reaction occurs:

   $^{233}U + ^{1}n \rightarrow ^{142}Ce + ^{82}Se + 10^{1}n$

   For every mole of ^{233}U that decays, 0.1355 g of mass is lost. How much energy is given off per mole of ^{233}U reacted?

2. In the spontaneous fission of plutonium-241, the following reaction occurs:

   $^{241}Pu + ^{1}n \rightarrow ^{104}Ru + ^{124}Sn + 14^{1}n$

   For every mole of ^{241}Pu that decays, 0.1326 g of mass is lost. How much energy is given off per mole of ^{241}Pu reacted?

3. The two rarer isotopes of hydrogen—deuterium and tritium—can also be fused to make helium by the following reaction:

   $^{2}H + ^{3}H \rightarrow ^{4}He + ^{1}n$

   In the course of this reaction, 0.01888 g of mass is lost. How much energy is emitted in the reaction of 1 mol of deuterium and tritium?

4. A process called *helium burning* is thought to occur inside older stars, forming carbon:

   $3^{4}He \rightarrow ^{12}C$

   If the reaction proceeds with 0.00781 g of mass lost on a molar basis, how much energy is given off?

5. Briefly describe how a nuclear reactor generates electricity.
6. Briefly describe the difference between how a nuclear reactor works and how a nuclear bomb works.

7. What is a chain reaction?
8. Why must uranium be enriched to supply nuclear energy?

Answers

1. 1.22×10^{13} J
2. 1.19×10^{13} J
3. 1.70×10^{12} J
4. 7.03×10^{11} J
5. A nuclear reactor generates heat, which is used to generate steam that turns a turbine to generate electricity.
6. Both nuclear reactor and nuclear bomb are powered by fission reaction however, in a nuclear reactor, the fission is monitored and controlled to occur continuously for a much longer time. In a nuclear bomb, the reaction is uncontrolled to explode in one event.
7. A chain reaction is an ever-expanding series of processes that, if left unchecked, can cause a runaway reaction and possibly an explosion.
8. Natural uranium ores contain only 0.7% U-235. Most nuclear reactors require enriched U-235 for their fuel.

11.6: Chapter Summary

Additional Exercises

1. Given that many elements are metals, suggest why it would be unsafe to have radioactive materials in contact with acids.
2. Many alpha-emitting radioactive substances are relatively safe to handle, but inhaling radioactive dust can be very dangerous. Why?
3. Uranium can be separated from its daughter isotope thorium by dissolving a sample in acid and adding sodium iodide, which precipitates thorium(III) iodide:

   $Th^{3+}(aq) + 3I^-(aq) \rightarrow ThI_3(s)$

   If 0.567 g of Th^{3+} were dissolved in solution, how many milliliters of 0.500 M NaI(aq) would have to be added to precipitate all the thorium?
4. Thorium oxide can be dissolved in an acidic solution:

   $ThO_2(s) + 4H^+ \rightarrow Th^{4+}(aq) + 2H_2O(\ell)$

   How many milliliters of 1.55 M HCl(aq) are needed to dissolve 10.65 g of ThO_2?
5. Radioactive strontium is dangerous because it can chemically replace calcium in the human body. The bones are particularly susceptible to radiation damage. Write the nuclear equation for the beta emission of strontium-90.
6. Write the nuclear equation for the beta emission of iodine-131, the isotope used to diagnose and treat thyroid problems.
7. A common uranium compound is uranyl nitrate hexahydrate [$UO_2(NO_3)_2$_6H_2O]. What is the formula mass of this compound?
8. Plutonium forms three oxides: PuO, PuO_2, and Pu_2O_3. What are the formula masses of these three compounds?
9. A banana contains 600 mg of potassium, 0.0117% of which is radioactive potassium-40. If 1 g of potassium-40 has an activity of 2.626×10^5 Bq, what is the activity of a banana?

10. Smoke detectors typically contain about 0.25 mg of americium-241 as part of the smoke detection mechanism. If the activity of 1 g of americium-241 is 1.26×10^{11} Bq, what is the activity of americium-241 in the smoke detector?
11. Uranium hexafluoride (UF_6) reacts with water to make uranyl fluoride (UO_2F_2) and hydrogen fluoride (HF). Balance the following chemical equation:

    $UF_6 + H_2O \rightarrow UO_2F_2 + HF$
12. The cyclopentadienyl anion ($C_5H_5^-$) is an organic ion that can make ionic compounds with positive ions of radioactive elements, such as Np^{3+}. Balance the following chemical equation:

    $NpCl_3 + Be(C_5H_5)_2 \rightarrow Np(C_5H_5)_3 + BeCl_2$

Answers

1. Acids can dissolve metals, making aqueous solutions.
2. Alpha rays are dangerous only when the alpha emitter is in direct contact with tissue cells inside the body.
3. 14.7 mL
4. 104 mL
5. $^{90}_{38}\mathrm{Sr} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{90}_{39}\mathrm{Y}$
6. $^{131}_{53}\mathrm{I} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{131}_{54}\mathrm{Xe}$
7. 502 g/mol
8. PuO = 260.06 g/mol; PuO_2 = 276.06 g/mol; Pu_2O_3 = 536.12 g/mol
9. about 18 Bq
10. 3.15×10^7 Bq
11. $UF_6 + 2H_2O \rightarrow UO_2F_2 + 4HF$
12. $2NpCl_3 + 3Be(C_5H_5)_2 \rightarrow 2Np(C_5H_5)_3 + 3BeCl_2$

5.S: Nuclear Chemistry (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

Some atoms have unstable nuclei that emit particles and high-energy electromagnetic radiation to form new elements that are more stable. This emission of particles and electromagnetic radiation is called **radioactivity**. There are three main types of spontaneous radioactive emission: **alpha particles**, which are equivalent to helium nuclei; **beta particles**, which are electrons; and **gamma radiation**, which is high-energy electromagnetic radiation. Another type of radioactive process is **spontaneous fission**, in which large nuclei spontaneously break apart into smaller nuclei and, often, neutrons. In all forms of radioactivity, new elements are formed from the radioactive reactants.

Radioactive isotopes decay at different rates. The rate of an isotope's decay is expressed as a **half-life**, which is the amount of time required for half of the original material to decay. The length of its half-life is a characteristic of the particular isotope and can range from less than microseconds to billions of years.

Amounts of radioactivity are measured in several different ways. A **becquerel** is equal to one radioactive decay per second. A **curie** represents 3.7×10^{10} decays per second. Other units describe the amount of energy absorbed by body tissues. One **rad** is equivalent to 0.01 joule of energy absorbed per gram of tissue. Different tissues react differently to different types of radioactivity. The **rem** unit takes into account not only the energy absorbed by the tissues, but also includes a numerical multiplication factor to account for the type of radioactivity and the type of tissue. The average annual radiation exposure of a person is less than 360 millirem, over 80% of which is from natural sources. Radioactivity can be detected using photographic film or other devices such as **Geiger counters**.

Radioactive isotopes have many useful applications. They can be used as **tracers** to follow the journey of a substance through a system, like an underground waterway or a metabolic pathway. Radioactive isotopes can be used to date objects, since the amount of parent and daughter isotopes can sometimes be measured very accurately. Radioactive emission can be used to sterilize food for a longer edible lifetime. There are also a number of diagnostic and therapeutic medical applications for radioactive isotopes.

Radioactive processes occur with simultaneous changes in energy. This **nuclear energy** can be used to generate power for human use. **Nuclear reactors** use the energy released by fission of large isotopes to generate electricity. When carefully controlled, fission can produce a **chain reaction** that facilitates the continuous production of energy. If not carefully controlled, a very quick production of energy can result, as in an **atomic bomb**. Natural uranium does not contain enough of the proper isotope of uranium to work in a nuclear reactor, so it must first be **enriched** in uranium-235. Forcing small nuclei together to make larger nuclei, a process called **fusion**, also gives off energy; however, scientists have yet to achieve a controlled fusion process.

CHAPTER OVERVIEW

6: Molecules and Compounds

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multiatom combinations are called moleculesThe smallest part of a substance that has the physical and chemical properties of that substance.. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

6.1: Sugar and Salt

Sodium chloride, also known as table salt, is an ionic compound with the chemical formula NaCl, representing a 1:1 ratio of sodium and chloride ions. It is commonly used as a condiment and food preservative. Salt can be created by adding two very reactive elements together: sodium ($\mathrm{Na(s)}$) metal and chlorine ($\mathrm{Cl_2(g)}$) gas.

$$2\,\mathrm{Na(s)} + \mathrm{Cl_2(g)} \rightarrow 2\,\mathrm{NaCl(s)} \tag{6.1.1}$$

The element sodium (Figure $6.1.1a$) is a very reactive metal; given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine (Figure $6.1.1b$) is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride (Figure $6.1.1c$), known simply as salt.

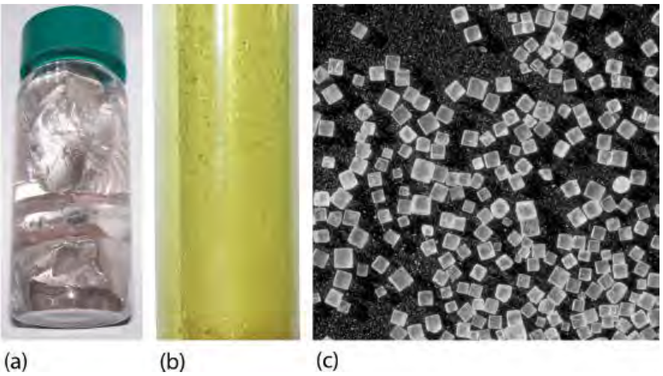


Figure $6.1.1$: Sodium + Chlorine = Sodium Chloride (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride—salt—which is necessary for our survival. Source: Photo on the left courtesy of reenhorn1 and photo in the center courtesy of Benjah-bmm27. Photo on the right © Thinkstock.

Salt is necessary for life. $\mathrm{Na^+}$ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. $\mathrm{Cl^-}$ ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavoring known, and one of the few rocks we eat. Clearly when the elemental sodium and chlorine combine (Equation 6.1.1), the resulting salt product has radically different properties (both physical and chemical). This reaction is spectacular to observe (Video $6.1.1$).

Video* $6.1.1$*: Making Table Salt using Sodium Metal and Chlorine gas

Another compound is sugar, which is the generic name for sweet, soluble carbohydrates, many of which are used in food. Sugar has the chemical formulate $\mathrm{C_{12}H_{22}O_{11}}$ and is constructed from different elements than salt: carbon, hydrogen and oxygen. While sugar qualitatively resembles table salt (often confused in the kitchen), the two have distinctly different physical and chemical properties. There are various types of sugar derived from different sources. While sugar is made with carbon, hydrogen, and oxygen, it is considerably harder to synthesize from its constituent elements than table salt is (Equation 6.1.1). However, the thermal decomposition is considerably easier and can be represented as a dehydration of sucrose to pure carbon and water vapor in Equation 6.1.2, and demonstrated in Video $6.1.2$.

$$C_{12}H_{22}O_{11}(s) + \text{heat} \rightarrow 12\,C(s) + 11\,H_2O(g) \tag{6.1.2}$$

Video 6.1.2: *A science experiment in the kitchen shows what happens to sugar molecules when they are heated. The experiment did not disappoint!*

As with salt, sugar has radically different properties (both physical and chemical) than its constituent elements. This difference in properties, of constituent elements and compounds, is a central feature of chemical reactions.

Contributions & Attributions

- Wikipedia

6.2: Compounds Display Constant Composition

When building a house, the starting point is a blueprint of what the house will look like. The plan states how many windows and what kind, how many doors and what style, how many rooms and what type (bedroom, kitchen, other). The blueprint shows how the different pieces will go together to make the house. As long as the blueprint is followed and exactly the same items are used, the house will be identical to its blueprint.

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances—carbon and hydrogen.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** is a change that produces matter with a different composition. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes into carbon and water. Water is still a compound, but one which cannot be broken down into hydrogen and oxygen by heating. Instead, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a white solid which is essential for all living things (see below).

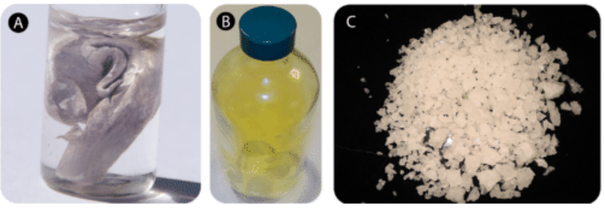


Figure 6.2.1: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine.

Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- A chemical change is a change that produces matter with a different composition.

6.3: Chemical Formulas- How to Represent Compounds

Learning Objectives

- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a 2:1 ratio. The chemical formula for water is H_2O. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is H_2SO_4.

Certain groups of atoms are bonded together to form what is called a polyatomic ion that acts as a single unit. Polyatomic ions are discussed in more detail in Section 5.5. Polyatomic ions are enclosed in parenthesis followed by a subscript if more than one of the same ion exist in a chemical formula. The formula $Ca_3(PO_4)_2$ represents a compound with the following:

3 Ca atoms + 2 PO_4^{3-} ions

To count the total number of atoms for formulas with polyatomic ions enclosed in parenthesis, use the subscript as a multiplier for each atom or number of atoms.

$$Ca_3(PO_4)_{2}$$

$$3\text{ Ca} + 2 \text{ x1 P} + 2 \text{ x 4 O} = 3 \text{ Ca atoms} + 2 \text{ P atoms} + 8 \text{ O atoms}$$

Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

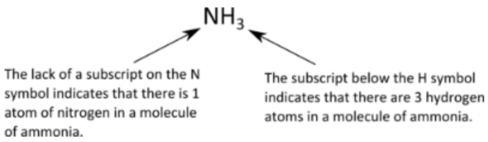


Figure 6.3.1: The molecular formula for ammonia. NH3. There is one atom of nitrogen and 3 atoms of hydrogen in a molecule of ammonia.

Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, which is not of concern at this point. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write N_3H for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula.

- molecular formula: $C_6H_{12}O_6$
- empirical formula: CH_2O

There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Summary

- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Polyatomic ions in chemical formulas are enclosed in parentheses followed by a subscript if more than one of the same type of polyatomic ion exist.
- Molecular formulas do not indicate how the atoms are arranged in the molecule.
- The empirical formula tells the lowest whole-number ratio of elements in a compound. The empirical formula does not show the actual number of atoms.

6.4: A Molecular View of Elements and Compounds

Learning Objectives

- Classify substances as atomic elements, molecular elements, molecular compounds, or ionic compounds.

Atomic Elements

Most elements exist with **individual atoms** as their basic unit. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

Molecular Elements

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called **molecules**. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Table 6.4.1: Elements That Exist as Diatomic Molecules

Hydrogen, H	Oxygen	Nitrogen	Fluorine	Chlorine	Bromine	Iodine

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules—a molecule with only two atoms (Table 6.4.1). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2, N_2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 (Figure 6.4.1).

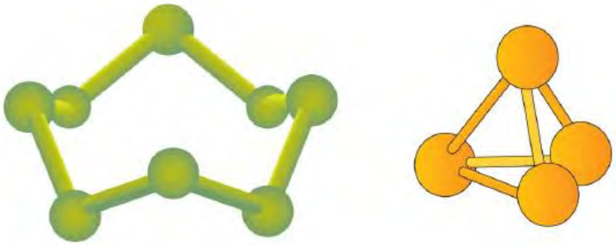


Figure 6.4.1: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 6.4.1 shows two examples of how molecules will be represented in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

Ionic Compounds

The elements in the periodic table are divided into specific groupings; the metals, the non-metals, the semi-metals, and so on. These groupings are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). The basic unit of ionic compounds is the **formula unit**.

Molecular Compounds

Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds. Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals) and silicon dioxide (SiO_2) will also be a covalently bound molecule (a semi-metal and a non-metal). The basic unit of molecular compounds is the **molecule**.

✓ Example 6.4.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. Fe
b. PCl_3
c. LiBr
d. P_4
e. oxygen gas

Solution

a. **Fe** (iron) is an element that is represented with no subscript, so it is an **atomic element**.
b. **PCl_3** is made up of two nonmetals, so it is a **molecular compound**.
c. **LiBr** is made up of lithium, a metal, and bromine, a nonmetal, so it is an **ionic compound**.
d. **P_4** is a substance that is made up of four atoms of the same element, so it is a **molecular element**.
e. The formula for **oxygen gas** is **O_2** so it is a molecular element.

? Exercise 6.4.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. I_2
b. He
c. H_2O
d. Al
e. CuCl

Answer a:

molecular element

Answer b:

atomic element

Answer c:

molecular compound

Answer d:

atomic element

Answer e:

ionic compound

6.5: Writing Formulas for Ionic Compounds

Learning Objectives

- Write the correct formula for an ionic compound.
- Recognize polyatomic ions in chemical formulas.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like $NaCl$, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charges to cancel each other out.

✓ Example 6.5.1: Aluminum Nitride and Lithium Oxide

Write the formulas for aluminum nitride and lithium oxide.

Solution

Solution to Example 5.5.1

	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Al^{3+} N^{3-}	Li^{+} O^{2-}
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

An alternative way to writing a correct formula for an ionic compound is to use the **crisscross method**. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped.

✓ Example 6.5.2: The Crisscross Method for Lead (IV) Oxide

Write the formula for lead (IV) oxide.

Solution

Solution to Example 5.5.2

Crisscross Method	Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Pb^{4+} O^{2-}

Crisscross Method	Write the formula for lead (IV) oxide
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

Exercise 6.5.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

> *Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.*

Example 6.5.3: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Solution to Example 5.5.3

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$Na^+ \quad S^{2-}$
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all subscripts that are 1.	Na_2S

Exercise 6.5.3

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a:

NaBr

Answer b:

LiCl

Answer c:

MgO

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Table $6.5.1$ lists the most common polyatomic ions.

Table $6.5.1$: Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

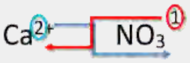
Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 6.5.4: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Solution to Example 5.5.4

Crisscross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$Ca^{2+} \quad NO_3^-$
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 6.5.5

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion.

Solution

Solution to Example 5.5.5

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 6.5.5

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

$MgCO_3$

Answer b:

$Al(CH_3COO)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds.

Method 1

Compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 [or 2A] metal) and a nonmetallic element (bromine, a group 17 [or 7A] nonmetal). Therefore, it is most likely an ionic compound (in fact, it *is* ionic). In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 [or 5A], and oxygen, from group 16 [or 6A]. It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Method 2

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

✓ Example 6.5.6

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. NH_4Cl
d. OF_2

Solution

Solution to Example 5.5.6

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic via method 1.	Na_2O, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic via method 1	PCl_3, not ionic
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic via method 2	NH_4Cl, ionic
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic via method 1	OF_2, not ionic

? Exercise 6.5.6

Identify each compound as ionic or not ionic.

a. N_2O
b. $FeCl_3$
c. $(NH_4)_3PO_4$
d. $SOCl_2$

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Summary

Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

6.6: Nomenclature- Naming Compounds

Nomenclature is the process of naming chemical compounds so that they can be easily identified as separate chemicals. The primary function of chemical nomenclature is to ensure that a spoken or written chemical name leaves no ambiguity concerning which chemical compound the name refers to—each chemical name should refer to a single substance. A less important aim is to ensure that each substance has a single name, although a limited number of alternative names is acceptable in some cases. Preferably, the name also conveys some information about the structure or chemistry of a compound. A common name will often suffice to identify a chemical compound in a particular set of circumstances. To be more generally applicable, the name should indicate at least the chemical formula. To be more specific still, the three-dimensional arrangement of the atoms may need to be specified.

Contributions & Attributions

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6.7: Naming Ionic Compounds

Learning Objectives

- To use the rules for naming ionic compounds.

After learning a few more details about the names of individual ions, you will be one step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system**, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 6.7.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 6.7.1 lists the elements that use the common system, along with their respective cation names.

Table 6.7.2: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion

Ion	Name
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix -*ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 6.7.2 lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

✓ Example 6.7.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. $SO_3{}^{2-}$
d. $NH_4{}^+$
e. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

? Exercise 6.7.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. $SO_4{}^{2-}$
d. Ba^{2+}
e. $HCO_3{}^-$

Answer a:

iron(II) ion

Answer b:

iron(III) ion

Answer c:

sulfate ion

Answer d:

barium ion

Answer e:

hydrogen carbonate ion or bicarbonate ion

✓ Example 6.7.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Solution

a. Br^{-}
b. PO_4^{3-}
c. Cu^{2+}
d. Mg^{2+}

? Exercise 6.7.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a:

F^{-}

Answer b:

CO_3^{2-}

Answer c:

Sn^{2+}

Answer d:

K^{+}

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 6.7.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

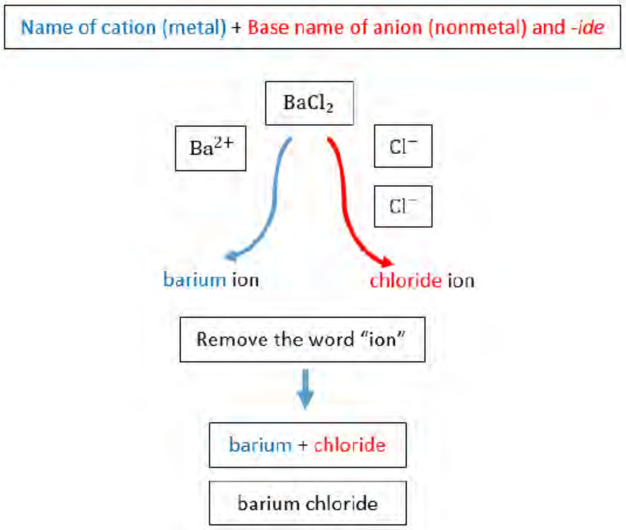


Figure 6.7.1: Naming $BaCl_2$
Naming formula: Name of metal cation + base
name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

✓ Example 6.7.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

? Exercise 6.7.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

Answer c:

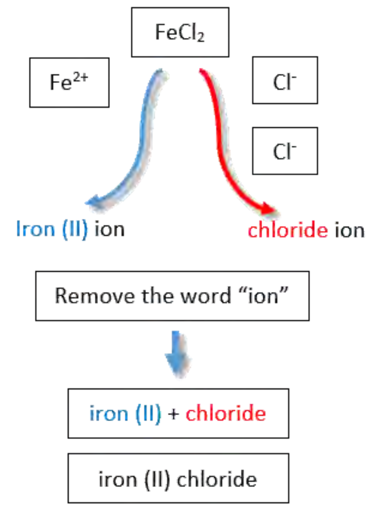
calcium phosphide

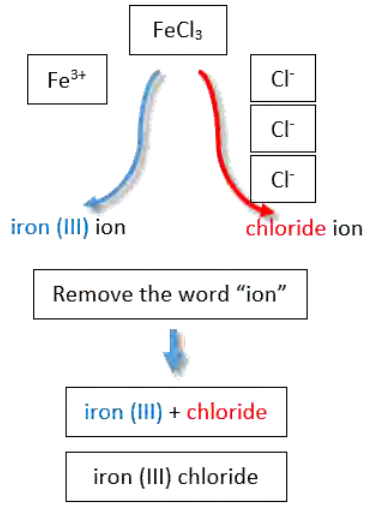
Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 6.7.2).

Table 6.7.3: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.

Name of cation (metal) + (Roman Numeral in parenthesis) + Base name of anion (nonmetal) and *-ide*



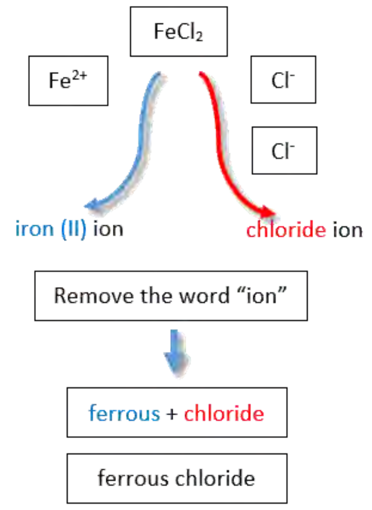


If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 6.7.3) .

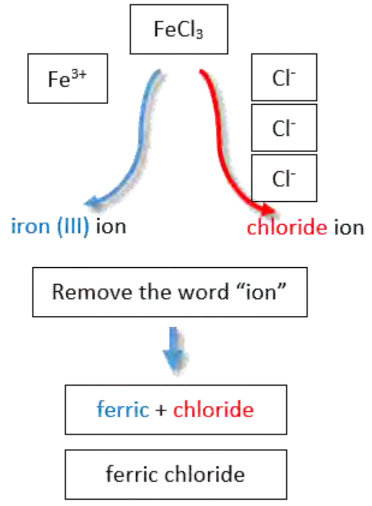
Table 6.7.4: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Old/Common System.

"Old" base name of cation (metal) and *-ic or -ous* + Base name of anion (nonmetal) and *-ide*

-ous (for ions with lower charge)



-ic (for ions with higher charge)



Example 6.7.4:

Name each ionic compound.

a. Co_2O_3
b. $FeCl2$

Solution

Solutions to Example 5.7.4

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloride always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−. • This means that the one iron ion must have a 2+ charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

? Exercise 6.7.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium sulfate compound in Figure 6.7.6.

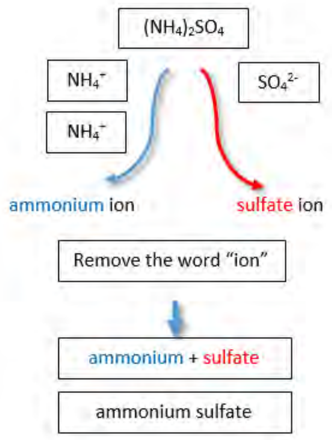


Figure 6.7.2: Naming Ionic Compounds with Polyatomic Ions

✓ Example 6.7.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Solutions to Example 5.7.5

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound's name is iron(II) phosphate.	iron(II) phosphate

? Exercise $6.7.5A$

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Figure $6.7.1$ is a synopsis of how to name simple ionic compounds.

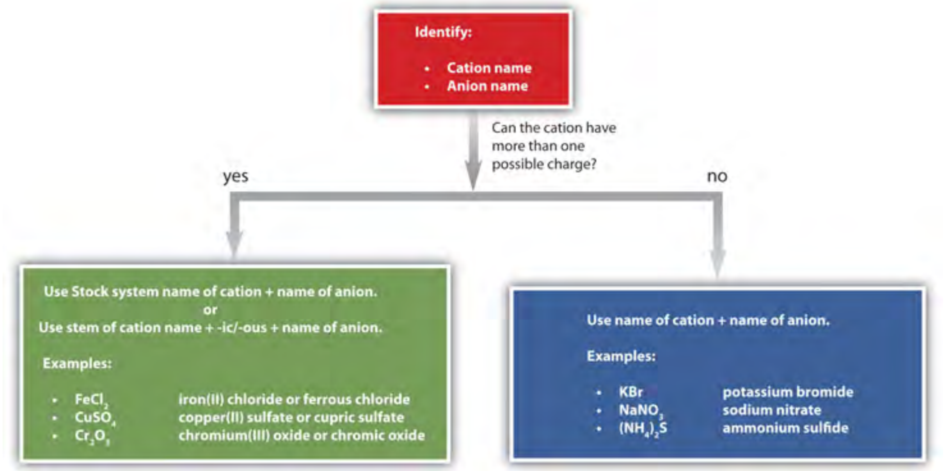


Figure $6.7.3$: A Guide to Naming Simple Ionic Compounds.

Exercise $6.7.5B$

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

ammonium phosphate

Answer d:

gold(III) fluoride or auric fluoride

Answer e:

silver fluoride

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parentheses.
- Ternary compounds are composed of three or more elements.

- **5.7: Naming Ionic Compounds** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

6.8: Naming Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride $(NaCl)$. Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **electrons** in such a way that a **bond** forms between a pair of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

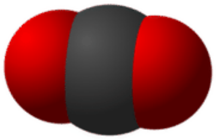


Figure 6.8.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.

Figure 6.8.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

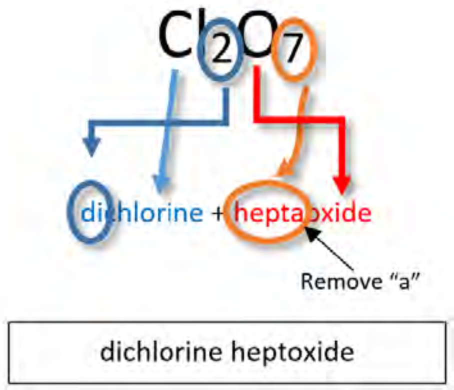
Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 6.8.1 lists these numerical prefixes.

Table 6.8.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

Number of Atoms in Compound	Prefix on the Name of the Element
7	hepta-
8	octa-
9	nona-
10	deca-

*This prefix is not used for the first element’s name.



Note

- Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetr**o**xide instead of tetr**ao**xide.
- The prefix is "mono" is not added to the first element’s name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table 6.8.2.

Table 6.8.2

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the *mono*- prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono*- and the *a* of *hepta*- are dropped from the name when paired with oxide.

Exercise 6.8.1

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

6.9: Naming Acids

A spot test for gold has been in use for decades. The sample is first treated with nitric acid. Other metals may react or dissolve in this acid, but gold will not. Then the sample is added to a mixture of nitric acid and hydrochloric acid. Gold will only dissolve in this mixture. The term "acid test" arose from the California gold rush in the late 1840's when this combination was used to test for the presence of real gold. It has since come to mean, "tested and approved" in a number of fields.

Acids

An **acid** can be defined in several ways. The most straightforward definition is that an acid is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions $\left(\mathrm{H}^+\right)$ when dissolved in water.

Figure 6.9.1: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

This is a different type of compound than the others we have seen so far. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$\mathrm{HCl} \rightarrow \mathrm{H}^+ + \mathrm{Cl}^-$$

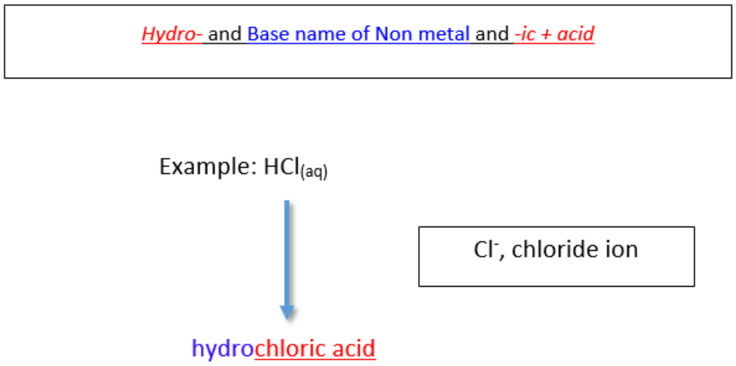
Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic.

Naming Binary acids (in aqueous form)

A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. The acid name begins with the prefix *hydro-*. followed by the base name of the anion, followed by the suffix *-ic*.



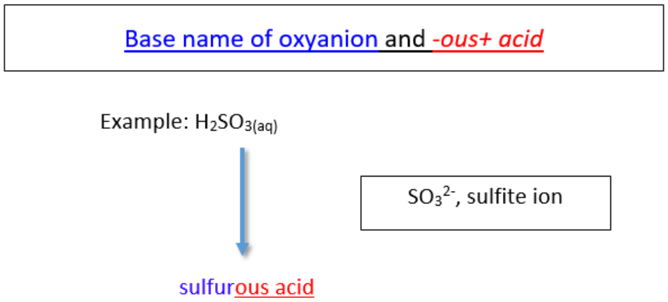
Formula for naming acids: Hydro- and Base name of nonmetal and -ic + acid. Example: HCl is hydrochloric acid.

Naming Oxyacids

An **oxyacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

a. Oxyanions with ***-ite*** ending.

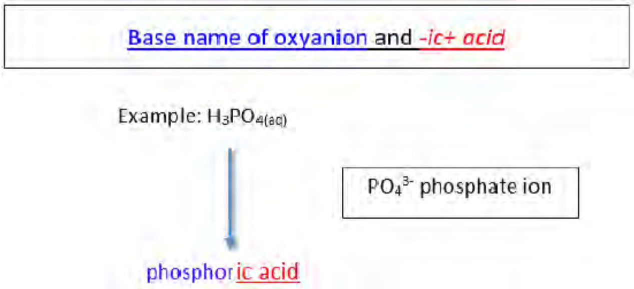
The name of the acid is the root of the anion followed by the suffix ***-ous***. There is no prefix.



Formula for naming oxyanions with -ite ending: Base name of oxyanion and -ous + acid. Example: H2SO3 is sulfurous acid.

b. Oxyanions with ***-ate*** ending.

The name of the acid is the root of the anion followed by the suffix ***-ic***. There is no prefix.



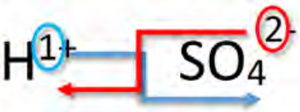
Formula for naming oxyanions with -ate ending: Base name of oxyanion and -ic + acid. Example: H3PO4 is phosphoric acid.

Note

The base name for sulfur containing oxyacid is ***sulfur-*** instead of just ***sulf-***. The same is true for a phosphorus containing oxyacid. The base name is ***phosphor-*** instead of simply ***phosph-***.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Another way to think about writing the correct formula is to utilize the crisscross method, shown below for sulfuric acid.



H2SO4: H has +1 charge and SO4 has -2 charge, so there must be 2 H+ and 1 SO4 for the charges to balance out.

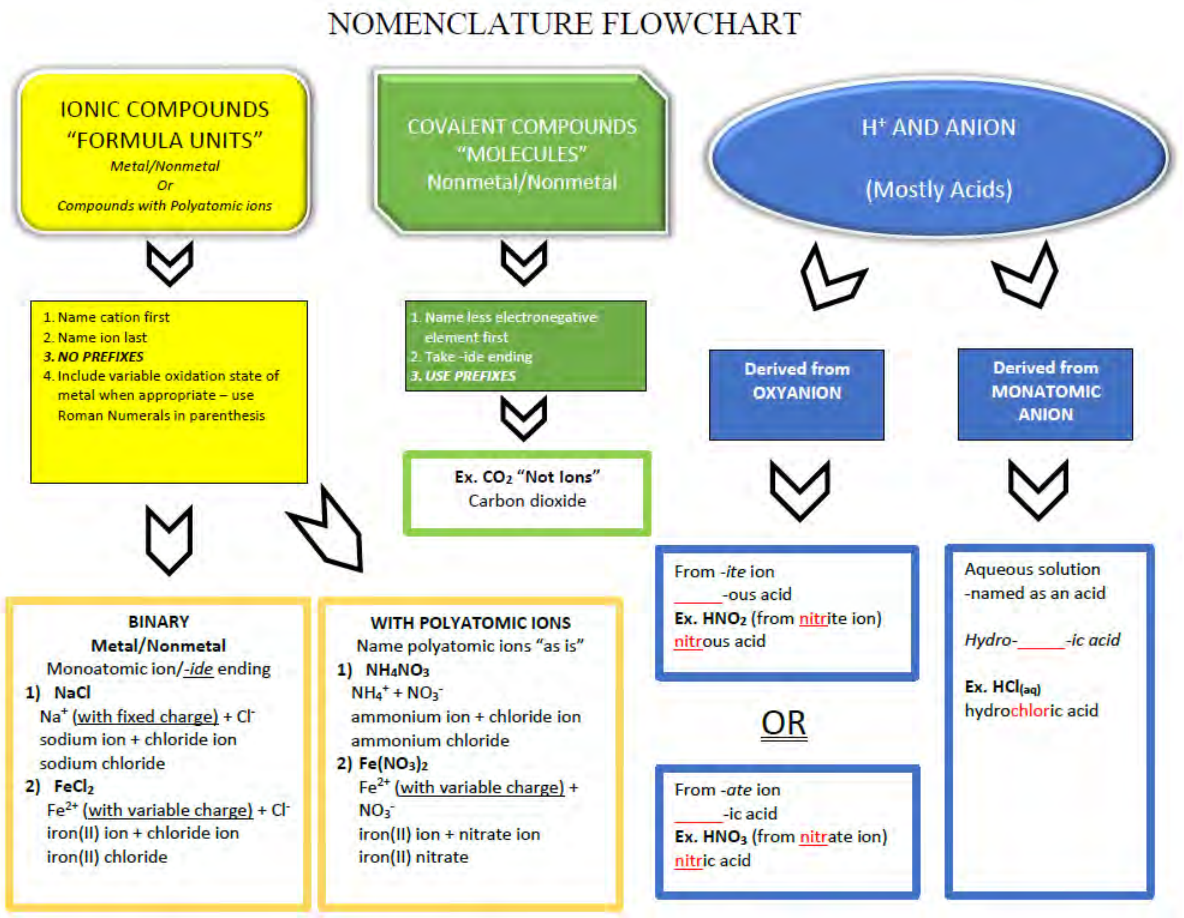
Formula: H_2SO_4

Figure 6.9.2: *Crisscross approach to writing formula for sulfuric acid.*

Summary

- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxyacids contain hydrogen, oxygen, and one other element.
- The name of the acid is based on the anion attached to the hydrogen.

6.10: Nomenclature Summary



NOMENCLATURE FLOWCHART

IONIC COMPOUNDS
"FORMULA UNITS"
Metal/Nonmetal
Or
Compounds with Polyatomic ions

1. Name cation first
2. Name ion last
3. ***NO PREFIXES***
4. Include variable oxidation state of metal when appropriate – use Roman Numerals in parenthesis

BINARY
Metal/Nonmetal
Monoatomic ion/*-ide* ending

1) **NaCl**
Na^+ (with fixed charge) + Cl^-
sodium ion + chloride ion
sodium chloride

2) **$FeCl_2$**
Fe^{2+} (with variable charge) + Cl^-
iron(II) ion + chloride ion
iron(II) chloride

WITH POLYATOMIC IONS
Name polyatomic ions "as is"

1) **NH_4NO_3**
NH_4^+ + NO_3^-
ammonium ion + chloride ion
ammonium chloride

2) **$Fe(NO_3)_2$**
Fe^{2+} (with variable charge) +
NO_3^-
iron(II) ion + nitrate ion
iron(II) nitrate

COVALENT COMPOUNDS
"MOLECULES"
Nonmetal/Nonmetal

1. Name less electronegative element first
2. Take -ide ending
3. ***USE PREFIXES***

Ex. CO_2 "Not Ions"
Carbon dioxide

H^+ AND ANION

(Mostly Acids)

Derived from OXYANION

From *-ite* ion
_____-ous acid
Ex. HNO_2 (from nitrite ion)
nitrous acid

OR

From *-ate* ion
_____-ic acid
Ex. HNO_3 (from nitrate ion)
nitric acid

Derived from MONATOMIC ANION

Aqueous solution
-named as an acid

Hydro-_____-ic acid

Ex. $HCl_{(aq)}$
hydrochloric acid

Contributions & Attributions

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6.11: Bonding Models and AIDS Drugs

6.12: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, or a Lewis diagram, or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H}\cdot$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\dot{\mathbf{H}} \quad or \quad \cdot\mathbf{H} \quad or \quad \underset{\cdot}{\mathbf{H}}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He}:$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1*s* subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li}\cdot$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be}:$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\dot{\mathbf{B}}:$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\cdot\dot{\mathbf{C}}:$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\cdot\underset{\cdot}{\dot{\mathbf{N}}}:$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that each side has no more than two electrons.

$$\cdot\underset{\cdot}{\ddot{\mathbf{O}}}:$$

Fluorine and neon have seven and eight dots, respectively:

$$:\underset{\cdot}{\ddot{\mathbf{F}}}:$$

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Ne}}}:$$

With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the $n = 3$ shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

✓ Example 6.12.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the $3s$ electrons:

$$\dot{Al}\,:$$

2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the $n = 4$ shell, there are six electrons. Its electron dot diagram is as follows:

$$\cdot\overset{\cdot}{\underset{\cdot\cdot}{\mathbf{Se}}}:$$

? Exercise 6.12.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a

$$\cdot\overset{\cdot}{\underset{\cdot}{\mathbf{P}}}:$$

Answer b

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Ar}}}:$$

Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis electron dot diagrams for ions have less (for cations) or more (for anions) dots than the corresponding atom.

6.13: Lewis Structures of Ionic Compounds- Electrons Transferred

Learning Objectives

- State the octet rule.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something: many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the *n* = 2 shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \curvearrowright \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

resulting in two ions—the Na^+ ion and the Cl^- ion:

$$\mathbf{Na}^+ \qquad \mathbf{:}\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^-$$

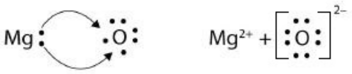
$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Na}^+ + \mathbf{:}\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^- \rightarrow Na^+Cl^- \quad or \quad NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

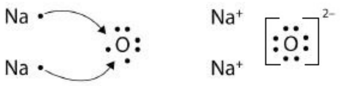
$$\mathbf{Mg}^{2+} + \left[:\ddot{\mathbf{O}}:\right]^{2-} \quad Mg^{2+}O^{2-} \text{ or } MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na}\cdot \curvearrowright \cdot\underset{\cdot}{\ddot{\mathbf{O}}}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



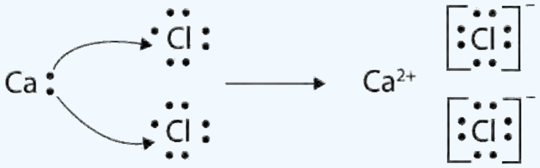
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na$_2$O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

✓ Example 6.13.1: Synthesis of Calcium Chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:

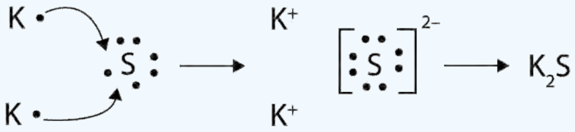


The oppositely charged ions attract each other to make CaCl$_2$.

? Exercise 6.13.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

Answer



Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

6.14: Covalent Lewis Structures- Electrons Shared

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won’t give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

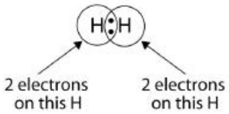
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H\cdot \quad \cdot H}$$

The two H atoms can share their electrons:

$$\mathbf{H:H}$$

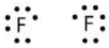
We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom’s valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity’s sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:

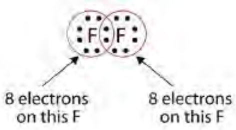
H–H

Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:

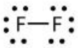


These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.

Note that each F atom has a complete octet around it now:



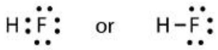
We can also write this using a dash to represent the shared electron pair:



There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:

The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

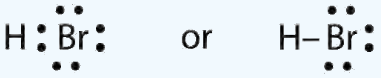
✓ Example 6.14.1:

Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:

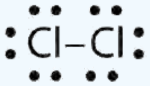
The two atoms can share their unpaired electron:



? Exercise 6.14.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer



When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far, and there are still some electrons that remain unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of C_2H_4. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two C atoms} = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

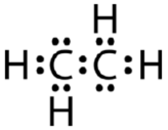


Figure 6.14.1: Incorrect dot structure of ethene. (CK12 License)

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

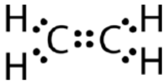


Figure 6.14.2: Correct dot structure for ethene. (CK12 License)

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

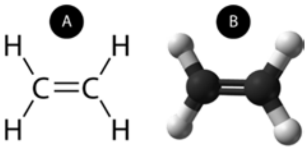


Figure 6.14.3: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

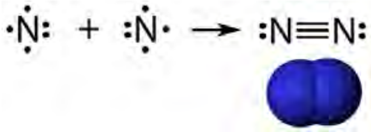


Figure 6.14.4: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributions & Attributions

-
- Anonymous by request

6.15: Writing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

For CO_3^{2-}, for example, we add two electrons to the total because of the −2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound’s structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

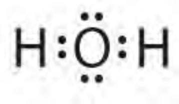
Now let’s apply this procedure to some particular compounds, beginning with one we have already discussed.

✓ Example 6.15.1: Water

Write the Lewis Structure for H_2O.

Solution

Solutions to Example 10.4.1

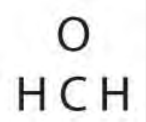
Steps for Writing Lewis Structures	Example 6.15.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of <u>8 valence electrons.</u>
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be <u>HOH.</u>
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

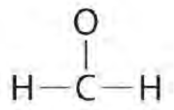
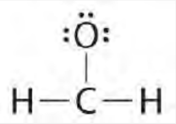
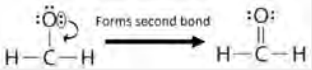
✓ Example 6.15.2

Write the Lewis structure for the CH_2O molecule

Solution

Solutions to Example 10.4.2

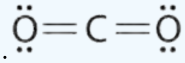
Steps for Writing Lewis Structures	Example 6.15.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = <u>12 valence electrons.</u>
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.

Steps for Writing Lewis Structures	Example 6.15.2
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 6.15.1

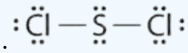
Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO_2



Carbon dioxide

Answer SCl_2



Sulfur dichloride

The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion $\left(H^+\right)$ attaches to the lone pair of an ammonia $\left(NH_3\right)$ molecule in a coordinate covalent bond.

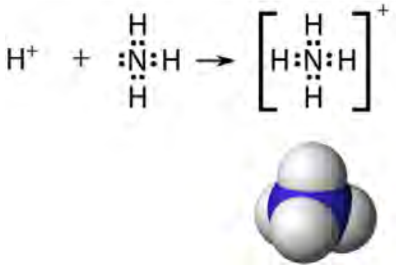


Figure 6.15.3: The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$$1 \text{ N atom} = 5 \text{ valence electrons}$$

$$4 \text{ H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{subtract 1 electron for the } 1+ \text{ charge of the ion}$$

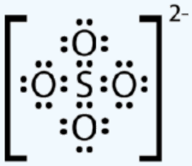
$$\text{total of 8 valence electrons in the ion}$$

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

✓ Exercise 6.15.2

Draw the Lewis electron dot structure for the sulfate ion.

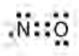
Answer (CK12 License)



Exceptions to the Octet Rule (BC Campus)

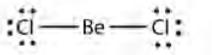
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

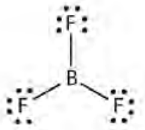


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

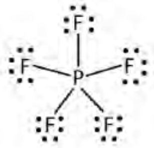
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

✓ Example 6.15.3: Octet Violations

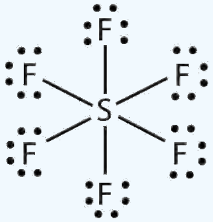
Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:

b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



? Exercise 6.15.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer

The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

6.16: Resonance - Equivalent Lewis Structures for the Same Molecule

Learning Objectives

- Explain the concept of resonance and how it works with within molecules.

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone (O_3) molecule in Figure $6.16.1$. There are a total of 18 electrons in the structure and so the following two structures are possible.

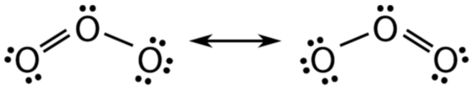


Figure 6.16.1: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left ($6.16.1$) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $O-O$ single bond and a double bond.

Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure $6.16.2$).

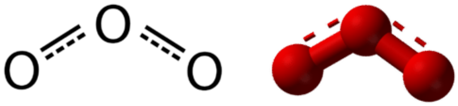


Figure 6.16.2: "Half-bond" model of ozone molecule. This is a better description of the electronic structure of ozone than either of the resonance structures in Figure 6.16.1.

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- in Figure $6.16.3$.

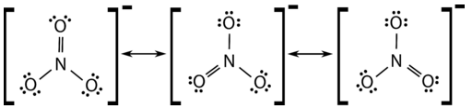


Figure 6.16.3: Resonance structure of nitrate anion.

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

- **10.6: Resonance - Equivalent Lewis Structures for the Same Molecule** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

6.17: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:

Figure 6.17.1: Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.

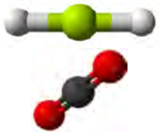


Figure 6.17.2: Beryllium hydride and carbon dioxide models. (CK12 Licence)

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

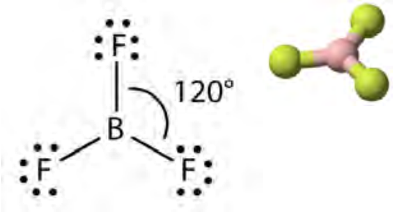


Figure 6.17.3: Boron trifluoride bonding. (CK12 Licence)

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

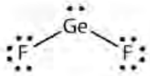


Figure 6.17.4: Germanium difluoride bonding.

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure $6.17.1$ Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

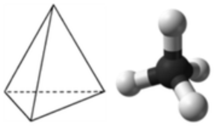


Figure 6.17.5: Tetrahedral structure of methane. (CK12 Licence)

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

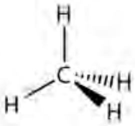


Figure 6.17.6: Methane bonding. (CK12 Licence)

NH_3 is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.

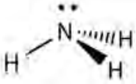


Figure 6.17.7: Ammonia bonding. (CK12 Licence)

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups, but only two of them are bonded to surrounding atoms.

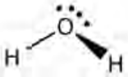


Figure 6.17.8: Water bonding.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom, is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 6.17.9.

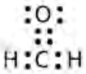


Figure 6.17.9: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

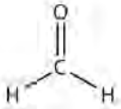


Figure 6.17.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table 6.17.1 summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table 6.17.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

✓ Example 6.17.1

What is the approximate shape of each molecule?

a. PCl_3
b. NOF

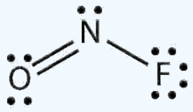
Solution

The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:

The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

? Exercise 6.17.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

? Exercise 6.17.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. (Hint: hydrogen is a terminal atom.)

Answer

Trigonal planar about both central C atoms.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

6.18: Electronegativity and Polarity- Why Oil and Water Don’t Mix

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom, and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur), and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure $6.18.1$.

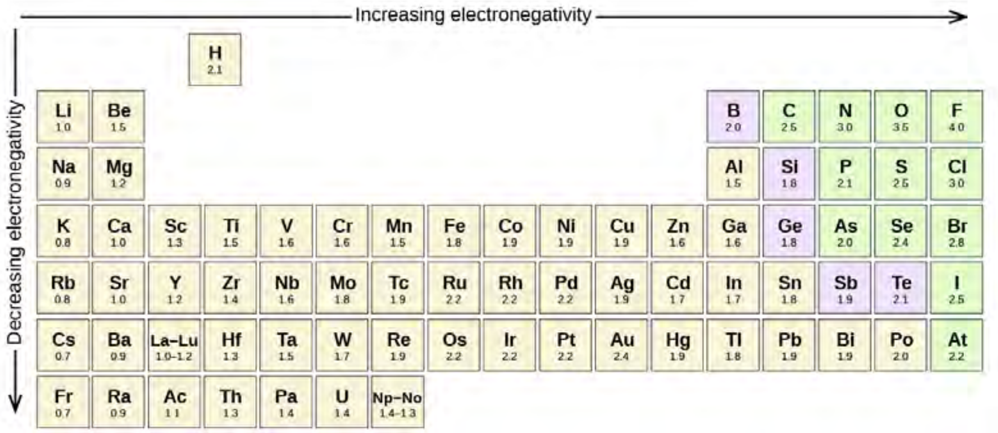


Figure $6.18.1$: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

difference of the electronegativities of the two atoms involved in the covalent bond

Electronegativity Difference	Bond Type
0–0.4	pure covalent
0.5–2.0	polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

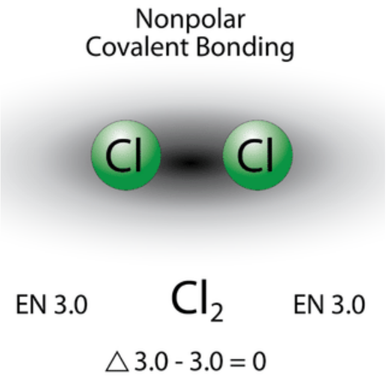


Figure 6.18.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathrm{Cl_2}$ molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta \mathrm{EN} = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

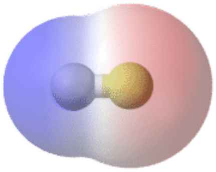


Figure 6.18.3: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

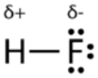


Figure 6.18.4: Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

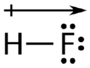


Figure 6.18.5: Use of crossed arrow to indicate polarity.

Electronegativity differences in bonding using Pauling scale. Differences in electronegativity classify bonds as covalent, polar covalent, or ionic.

Example 6.18.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 6.18.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 − 2.1 = 0.4. Thus we predict that this bond will be nonpolar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 − 2.1 = 1.4, so we predict that this bond will be polar covalent.

Exercise 6.18.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is generally useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical—bonded to the same element with no unshared pairs of electrons. **Polar molecules** are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well, as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming that you do, you can look at the structure of each one and decide if it is polar or not, *whether or not you know the* *individual atom's electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.

Figure 6.18.6: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $\left(\mathrm{CO_2}\right)$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

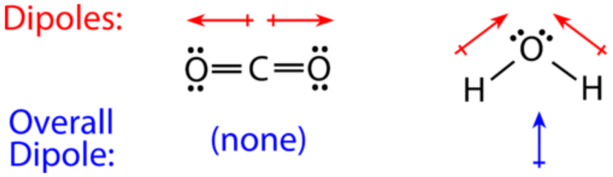


Figure 6.18.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $\left(\mathrm{BF_3}\right)$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $\left(\mathrm{NH_3}\right)$ is polar.

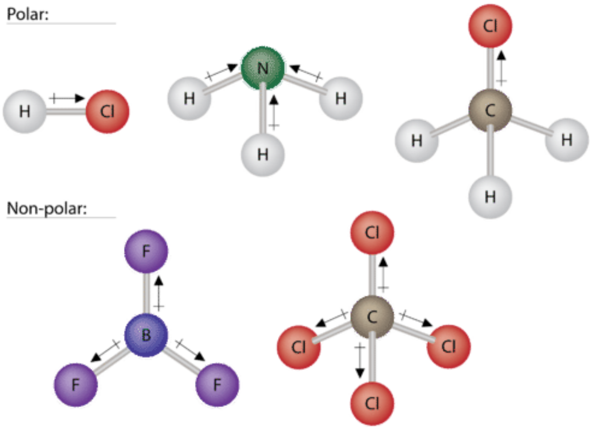


Figure 6.18.8: Some examples of polar and nonpolar molecules based on molecular geometry.

To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment do not cancel.

Steps to Identify Polar Molecules

1. Draw the Lewis structure.
2. Figure out the geometry (using VSEPR theory).
3. Visualize or draw the geometry.
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it).
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure 6.18.14). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

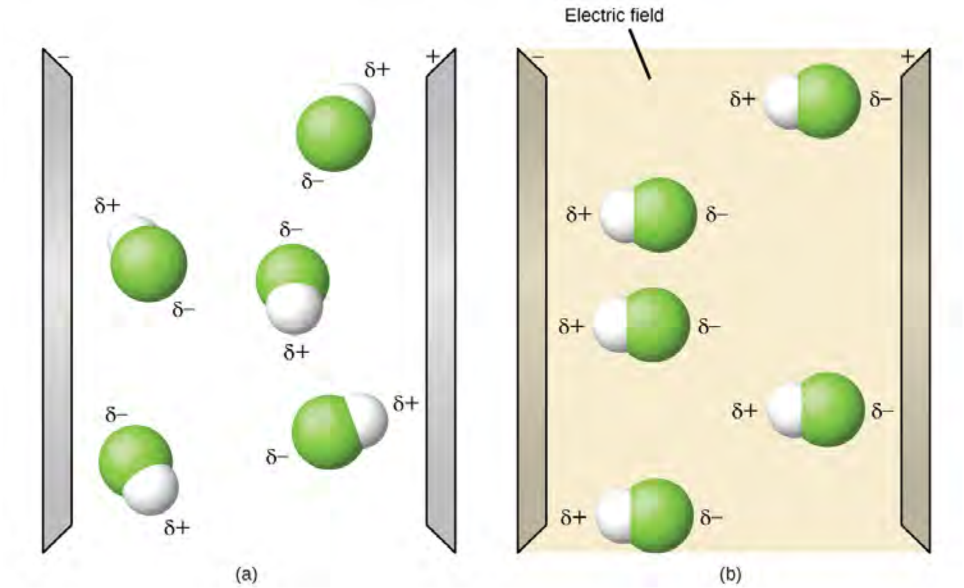


Figure 6.18.9: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

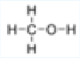
While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Example 6.18.2:

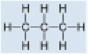
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: $H{-}C{\equiv}N$

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

? Exercise 6.18.2

Label each of the following as polar or nonpolar.

a. $\mathrm{SO_3}$
b. $\mathrm{NH_3}$

Answer a

nonpolar

Answer b

polar

Contributions & Attributions

- StackExchange (thomij).

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CHAPTER OVERVIEW

7: Stoichiometry

So far, we have talked about chemical reactions in terms of individual atoms and molecules. Although this works, most of the reactions occurring around us involve much larger amounts of chemicals. Even a tiny sample of a substance will contain millions, billions, or a hundred billion billions of atoms and molecules. How do we compare amounts of substances to each other in chemical terms when it is so difficult to count to a hundred billion billion? Actually, there are ways to do this, which we will explore in this chapter. In doing so, we will increase our understanding of stoichiometry, which is the study of the numerical relationships between the reactants and the products in a balanced chemical reaction.

7.1: Grade School Volcanoes, Automobiles, and Laundry Detergents

7.2: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

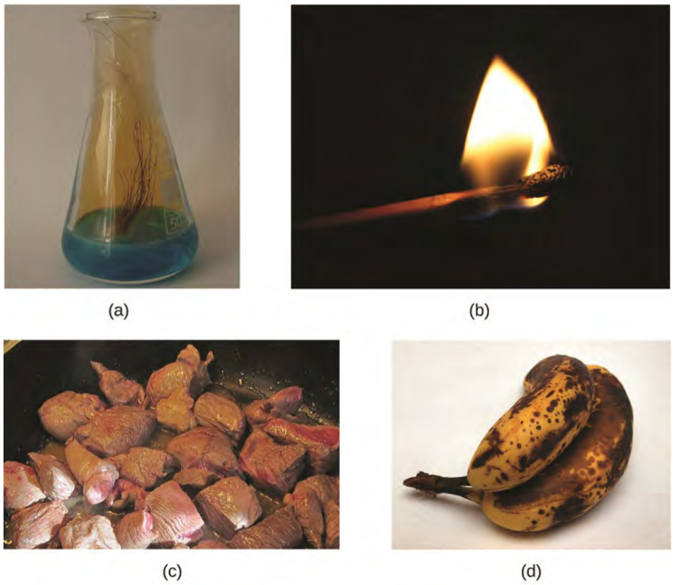

(a) (b) (c) (d)

Figure 7.2.1: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (Credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo.)

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include: reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure 7.2.1).

Video7.2.1: Evidence of a Chemical Reaction

✓ Example 7.2.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

a. Freezing liquid mercury.
b. Adding yellow to blue to make green.
c. Cutting a piece of paper into two pieces.
d. Dropping a sliced orange into a vat of sodium dydroxide.
e. Filling a balloon with natural air.

Solution

A, B, C, & E involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is D.

? Exercise 7.2.1

Which of the following is a chemical reaction?

a. Painting a wall blue.
b. A bicycle rusting.
c. Ice cream melting.
d. Scratching a key across a desk.
e. Making a sand castle.

Answer

B

✓ Example 7.2.2: Evidence of a Chemical Reaction

Which of the following is not a chemical reaction?

a. Shattering glass with a baseball.
b. Corroding metal.
c. Fireworks exploding.
d. Lighting a match.
e. Baking a cake.

Solution

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is A.

? Exercise 7.2.2

Which of the following is NOT a chemical reaction?

a. Frying an egg.
b. Slicing carrots.
c. A Macbook falling out of a window.
d. Creating ATP in the human body.
e. Dropping a fizzy tablet into a glass of water.

Answer

B and C

Summary

Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties.

7.3: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that—what is produced—or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\text{Flour} + \text{Butter} + \text{Sugar} + \text{Vanilla} + \text{Baking Soda} + \text{Eggs} + \text{Chocolate Chips}}_{\text{Ingredients} = \text{Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\mathrm{SO_2} + \mathrm{O_2}$, are reactants and sulfur trioxide, $\mathrm{SO_3}$, is the product.

$$\underbrace{2\,\mathrm{SO_2(g)} + \mathrm{O_2(g)}}_{\text{Reactants}} \rightarrow \underbrace{2\,\mathrm{SO_3(g)}}_{\text{Products}}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants} \rightarrow \text{Products}$$

There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table $7.3.1$ shows a listing of symbols used in chemical equations.

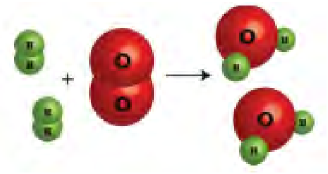
Table $7.3.1$: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)

Symbol	Description	Symbol	Description
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

$$2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{H_2O}\,(g)$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$\mathrm{Ca(NO_3)_2}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Ca(OH)_2}\,(s) + 2\mathrm{NaNO_3}\,(aq)$$

How much easier is that to read? Let's try it in reverse. Look at the following reaction in shorthand and write the word equation for the reaction:

$$\mathrm{Cu}\,(s) + \mathrm{AgNO_3}\,(aq) \rightarrow \mathrm{Cu(NO_3)_2}\,(aq) + \mathrm{Ag}\,(s)$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).
3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table 7.3.1: *Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

✓ Example 7.3.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
b. Gaseous propane, $\mathrm{C_3H_8}$, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane $(\mathrm{C_3H_8})$ and oxygen $(\mathrm{O_2})$

Product: carbon dioxide $(\mathrm{CO_2})$ and water $(\mathrm{H_2O})$

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g)$$

? Exercise 7.3.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.
b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{LiCl}\,(aq) + \mathrm{H_2O}\,(l)$
c. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

Answer b

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

Answer c

$Cu(s) + O_2(g) \rightarrow CuO(s)$

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

7.4: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products; and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

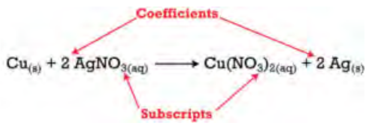
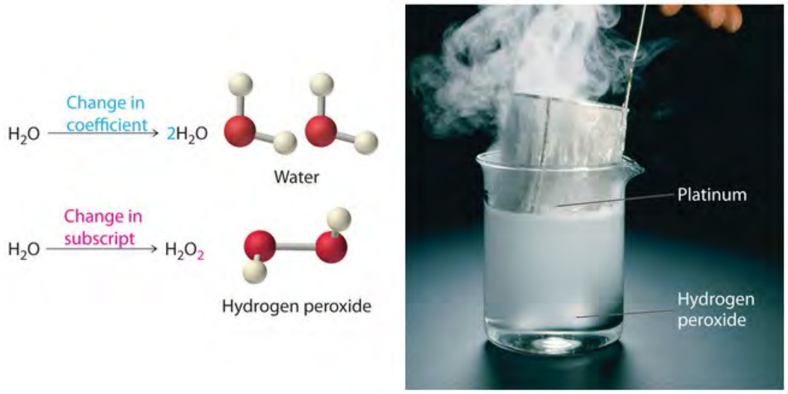


Figure 7.4.1: Balancing Equations. *You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.*

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 7.4.1.



Original molecule H2O: if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make H2O2, that's hydrogen peroxide.

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

Example 7.4.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).

$$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)}$$

Solution

Solutions to Example 7.4.1

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow \underline{7}CO_2(g) + H_2O(g)}$$ - 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow 7\,CO_2(g) + \underline{8}H_2O(g)}$$ - 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $$\mathrm{C_7H_{16}(l) + \underline{11}O_2(g) \rightarrow 7\,CO_2(g) + 8\,H_2O(g)}$$ - 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

✓ Example 7.4.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Solutions to Example 7.4.2

Steps	**Example**
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $C_8H_{18}(l) + \underline{\frac{25}{2}}O_2(g) \rightarrow 8\,CO_2(g) + 9\,H_2O(g)$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

✓ Example 7.4.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Solutions to Example 7.4.3

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $Pb(NO_3)_2(aq) + \underline{2}NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on reactant side, 1 Na atom on product side • 2 Cl atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow \underline{2}NaNO_3(aq) + PbCl_2(s)$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides
4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow 2\,NaNO_3(aq) + PbCl_2(s)$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides

? Exercise 7.4.1

Is each chemical equation balanced?

a. $2\,Hg(\ell)^{+}O_2(g) \rightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)$
c. $Mg(NO_3)_2(s) + 2\,Li(s) \rightarrow Mg(s)^{+}2\,LiNO_3(s)$

Answer a

yes

Answer b

no

Answer c

yes

? Exercise 7.4.2

Balance the following chemical equations.

a. $N_2(g) + O_2(g) \rightarrow NO_2(g)$
b. $Pb(NO_3)_2(aq) + FeCl_3(aq) \rightarrow Fe(NO_3)_3(aq) + PbCl_2(s)$
c. $C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Answer a

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$

Answer b

$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

Answer c

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- **Chemical reaction -** The process in which one or more substances are changed into one or more new substances.
- **Reactants -** The starting materials in a reaction.
- **Products -** Materials present at the end of a reaction.
- **Balanced chemical equation -** A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts -** Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient -** A small whole number that appears in front of a formula in a balanced chemical equation.

7.5: Climate Change - Too Much Carbon Dioxide

Carbon dioxide (CO_2) is an important heat-trapping (greenhouse) gas, which is released through human activities such as deforestation and burning fossil fuels, as well as natural processes such as respiration and volcanic eruptions. Figure 7.5.1 shows CO_2 levels during the last three glacial cycles, as reconstructed from ice cores.

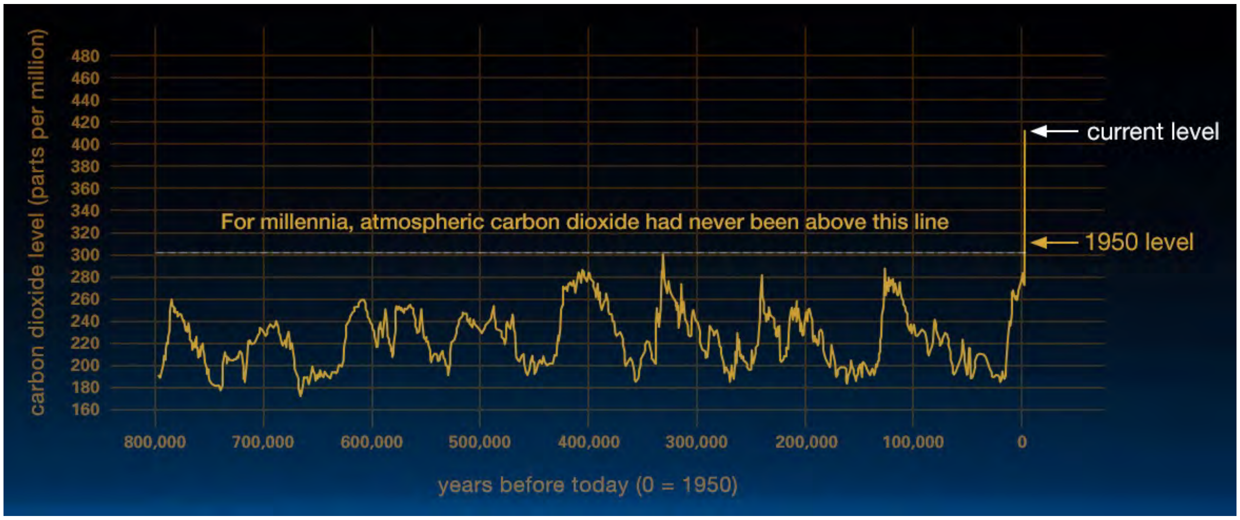


Figure 7.5.1: This graph, based on the comparison of atmospheric samples contained in ice cores and more recent direct measurements, provides evidence that atmospheric CO2 has increased since the Industrial Revolution. (Luthi, D., et al. 2008; Etheridge, D.M., et al. 2010; Vostok ice core data/J.R.R. Petit et al.; NOAA Mauna Loa CO2 record. NASA.) Line graph of carbon dioxide levels from 800 thousand before up until present day; current carbon dioxide level is higher than all historical levels.

Carbon dioxide (CO_2) is the primary greenhouse gas emitted through human activities. In 2015, CO_2 accounted for about 82.2% of all U.S. greenhouse gas emissions from human activities. Carbon dioxide is naturally present in the atmosphere as part of the Earth's carbon cycle (the natural circulation of carbon among the atmosphere, oceans, soil, plants, and animals). Human activities are altering the carbon cycle, both by adding more CO_2 to the atmosphere and by influencing the ability of natural sinks, like forests, to remove CO_2 from the atmosphere. While CO_2 emissions come from a variety of natural sources, human-related emissions are responsible for the increase that has occurred in the atmosphere since the industrial revolution.

The main human activity that emits CO_2 is the combustion of fossil fuels (coal, natural gas, and oil) for energy and transportation, although certain industrial processes and land-use changes also emit CO_2. As an example of how CO_2 can be generated, consider the combustion of octane, a component of gasoline:

$$2\,C_8H_{18}(l) + 21\,O_2(g) \rightarrow 16\,CO_2(g) + 18\,H_2O(g) \tag{7.5.1}$$

The balanced reaction in Equation 7.5.1 demonstrates that for every two molecules of octane that are burned, 16 molecules of CO_2 are generated.

Contributions & Attributions

- Earth Science Communications Team at NASA's Jet Propulsion Laboratory, California Institute of Technology
- EPA

7.6: Stoichiometry

Learning Objectives

- Explain the meaning of the term "stoichiometry".
- Determine the relative amounts of each substance in chemical equations.

You have learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this section you will explore the quantitative relationships that exist between the quantities of reactants and products in a balanced equation. This is known as **stoichiometry**.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry is actually Greek from two words: $\sigma\tau o\iota\kappa\eta\iota o\nu$, which means "element", and \(\mu \epsilon \tau \rho \omicron \nu), which means "measure".

Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

$$2\mathrm{CuSO_4} + 4\mathrm{KI} \rightarrow 2\mathrm{CuI} + 2\mathrm{K_2SO_4} + \mathrm{I_2}$$

The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 2 units of copper (II) sulfate $\left(\mathrm{CuSO_4}\right)$ we have, we need to have 4 units of potassium iodide (KI). For every two dozen copper (II) sulfates, we need 4 dozen potassium iodides. Because the unit "mole" is also a counting unit, we can interpret this equation in terms of moles, as well: For every two moles of copper (II) sulfate, we need 4 moles potassium iodide.

The production of ammonia $\left(\mathrm{NH_3}\right)$ from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightarrow 2\mathrm{NH_3}\,(g)$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

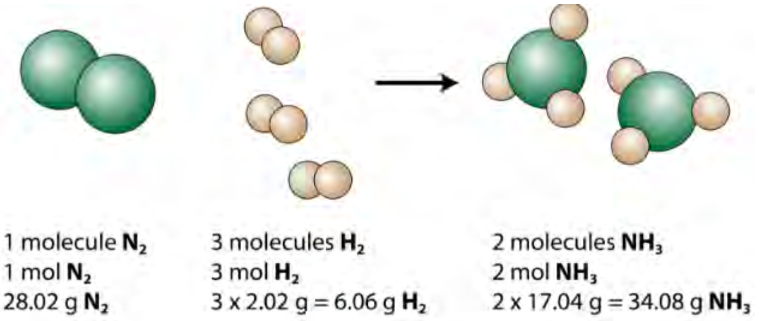


Figure 7.6.1: This representation of the production of ammonia from nitrogen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

We see that 1 molecule of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia. This is the smallest possible relative amount of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

The most useful quantity for counting particles is the mole. So if each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation.

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. $1\ \mathrm{mol}$ of nitrogen has a mass of $28.02\ \mathrm{g}$ while $3\ \mathrm{mol}$ of hydrogen has a mass of $6.06\ \mathrm{g}$ and $2\ \mathrm{mol}$ of ammonia has a mass of $34.08\ \mathrm{g}$

$$28.02 \text{ g } N_2 + 6.06 \text{ g } H_2 \rightarrow 34.08 \text{ g } NH_3$$

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.

✓ Example 7.6.1

The equation for the combustion of ethane (C_2H_6) is

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

a. Indicate the number of formula units or molecules in the balanced equation.
b. Indicate the number of moles present in the balanced equation.

Solution

a. Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.
b. Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.

? Exercise 7.6.1

For the following equation below, indicate the number of formula units or molecules, and the number of moles present in the balanced equation.

$$KBrO_3 + 6KI + 6HBr \rightarrow 7KBr + 3H_2O$$

Answer

One molecules of $KBrO_3$ plus six molecules of KI plus six molecules of HBr yields seven molecules of KBr plus three molecules of I_2 and three molecules of H_2O. One mole of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 plus three moles of H_2O.

Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- The coefficients of a balanced equation can be used to determine the ratio of moles of all substances in the reaction.

Vocabulary

- **Stoichiometry -** The calculation of quantitative relationships of the reactants and products in a balanced chemical equation.
- **Formula unit -** The empirical formula of an ionic compound.
- **Mole ratio -** The ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation.

7.7: Mole-to-Mole Conversions

/;Learning Objectives

- Use a balanced chemical equation to determine molar relationships between substances.

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). As follows, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4H_2 + 2O_2 \rightarrow 4H_2O$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22H_2 + 11O_2 \rightarrow 22H_2O$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ \mathrm{H_2} + 6.022 \times 10^{23}\ \mathrm{O_2} \rightarrow 12.044 \times 10^{23}\ \mathrm{H_2O}$$

These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ mol\ \mathrm{H_2} + 1\ mol\ \mathrm{O_2} \rightarrow 2\ mol\ \mathrm{H_2O}$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

$$2\,\mathrm{H_2} + \mathrm{O_2} \rightarrow 2\,\mathrm{H_2O}$$

Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level, but also in terms of molar amounts of reactants and products.* Thus, we can read this reaction as “two moles of hydrogen react with one mole of oxygen to produce two moles of water.”

By the same token, the ratios we constructed to describe a molecular reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\text{ mol }\mathrm{H_2}}{1\text{ mol }\mathrm{O_2}} \text{ or } \frac{1\text{ mol }\mathrm{O_2}}{2\text{ mol }\mathrm{H_2}}$$

$$\frac{2\text{ mol }\mathrm{H_2O}}{1\text{ mol }\mathrm{O_2}} \text{ or } \frac{1\text{ mol }\mathrm{O_2}}{2\text{ mol }\mathrm{H_2O}}$$

$$\frac{2\text{ mol }\mathrm{H_2}}{2\text{ mol }\mathrm{H_2O}} \text{ or } \frac{2\text{ mol }\mathrm{H_2O}}{2\text{ mol }\mathrm{H_2}}$$

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called *stoichiometry*.

Example 7.7.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?

Solution

Solutions to Example 8.3.1

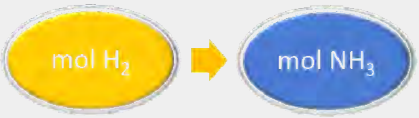
Steps for Problem Solving	How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?
Find a balanced equation that describes the reaction.	Unbalanced: $H_2 + O_2 \rightarrow H_2O$ Balanced: $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$
Identify the "given" information and what the problem is asking you to "find."	Given: moles H_2O Find: moles oxygen
List other known quantities.	1 mol O_2 = 2 mol H_2O
Prepare a concept map and use the proper conversion factor.	 $$\frac{1\ mol\ O_2}{2\ mol\ H_2O}$$
Cancel units and calculate.	$27.6\ \text{mol}\ \cancel{\text{H}_2\text{O}} \times \frac{1\ \text{mol}\ \text{O}_2}{2\ \text{mol}\ \cancel{\text{H}_2\text{O}}} = 13.8\ \text{mol}\ \text{O}_2$ To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

Example 7.7.2

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Solutions to Example 8.3.2

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Find a balanced equation that describes the reaction.	Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Balanced: $N_2 + \underline{3}H_2 \rightarrow \underline{2}NH_3$
Identify the "given" information and what the problem is asking you to "find."	Given: $H_2 = 4.20$ mol Find: mol of NH_3
List other known quantities.	3 mol H_2 = 2 mol NH_3
Prepare a concept map and use the proper conversion factor.	 $$\frac{2\ mol\ NH_3}{3\ mol\ H_2}$$

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Cancel units and calculate.	$4.20 \text{ mol } \cancel{H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } \cancel{H_2}} = 2.80 \text{ mol } NH_3$ The reaction of 4.20 mol of hydrogen with excess nitrogen produces 2.80 mol of ammonia.
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

? Exercise 7.7.3

a. Given the following balanced chemical equation:

$$\mathrm{C_5H_{12} + 8\,O_2 \rightarrow 5\,CO_2 + 6\,H_2O}$$

, How many moles of $\mathrm{H_2O}$ can be formed if 0.0652 mol of $\mathrm{C_5H_{12}}$ were to react?

b. Balance the following unbalanced equation and determine how many moles of $\mathrm{H_2O}$ are produced when 1.65 mol of NH_3 react:

$$\mathrm{NH_3 + O_2 \rightarrow N_2 + H_2O}$$

Answer a

0.391 mol H_2O

Answer b

$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol H_2O

Summary

- The balanced chemical reaction can be used to determine molar relationships between substances.

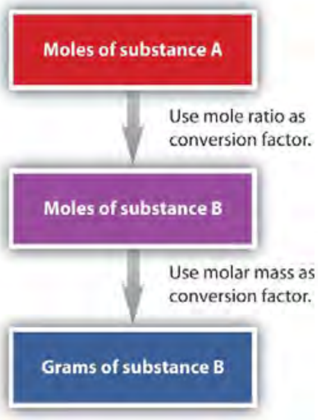
7.8: Making Molecules- Mole to Mass (or vice versa) and Mass-to-Mass Conversions

Learning Objectives

- Convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

Mole to Mass Conversions

We have established that a balanced chemical equation is balanced in terms of moles, as well as atoms or molecules. We have used balanced equations to set up ratios, in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions—such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that relation to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations. Flowchart of mole mass calculations: To convert from moles of substance A to moles of substance B, use the mole ratio conversion factor, and to convert from moles to grams of substance B, use molar mass conversion factor

As an example, consider the balanced chemical equation

$$\mathrm{Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3} \tag{7.8.1}$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then, using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

As usual, we start with the quantity we were given:

$$3.59\ \ \mathrm{mol}\ \cancel{\mathrm{Fe_2O_3}} \times \left(\frac{3\ \mathrm{mol\ SO_3}}{1\ \ \mathrm{mol}\ \cancel{\mathrm{Fe_2O_3}}}\right) = 10.77\ \mathrm{mol\ SO_3} \tag{7.8.2}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77\ \ \cancel{\mathrm{mol\ SO_3}} \times \left(\frac{80.06\ \mathrm{g\ SO_3}}{1\ \ \cancel{\mathrm{mol\ SO_3}}}\right) = 862\ \mathrm{g\ SO_3} \tag{7.8.3}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

> The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:
>
> $$3.59\ \ mol\ \cancel{Fe_2O_3} \times \underbrace{\left(\frac{3\ \ \cancel{mol\ SO_3}}{1\ \ mol\ \cancel{Fe_2O_3}}\right)}_{\text{converts to moles of SO}_3} \times \underbrace{\left(\frac{80.06\ g\ SO_3}{1\ \ \cancel{mol\ SO_3}}\right)}_{\text{converts to grams of SO}_3} = 862\ g\ SO_3$$
>
> We get exactly the same answer when combining all math steps together.

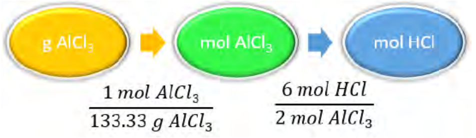
Example 7.8.1: Generation of Aluminum Oxide

How many moles of HCl will be produced when 249 g of $\mathrm{AlCl_3}$ are reacted according to this chemical equation?

$$2\,\mathrm{AlCl_3} + 3\,\mathrm{H_2O}(\ell) \rightarrow \mathrm{Al_2O_3} + 6\,\mathrm{HCl(g)}$$

Solution

Solutions to Example 8.5.1

Steps for Problem Solving	Example $7.8.1$
Identify the "given" information and what the problem is asking you to "find."	Given: 249 g AlCl3 Find: moles HCl
List other known quantities.	1 mol AlCl3 = 133.33 g AlCl3 6 mol of HCl to 2 mol $AlCl_3$
Prepare a concept map and use the proper conversion factor.	 Flowchart of needed conversion factors: 1 mole AlCl3 to 133.33 grams AlCl3, and 6 moles HCl to 2 moles AlCl3
Cancel units and calculate.	$249\;g\,\cancel{AlCl_3} \times \frac{1\;mol\,\cancel{AlCl_3}}{133.33\;g\,\cancel{AlCl_3}} \times \frac{6\,mol\,HCl}{2\;mol\,\cancel{AlCl_3}} = 5.60\,mol\,HCl$
Think about your result.	Since 249 g of $AlCl_3$ is less than 266.66 g, the mass for 2 moles of $AlCl_3$ and the relationship is 6 mol of HCl to 2 mol $AlCl_3$, the answer should be less than 6 moles of HCl.

Exercise 7.8.1: Generation of Aluminum Oxide

How many moles of $\mathrm{Al_2O_3}$ will be produced when 23.9 g of $\mathrm{H_2O}$ are reacted according to this chemical equation?

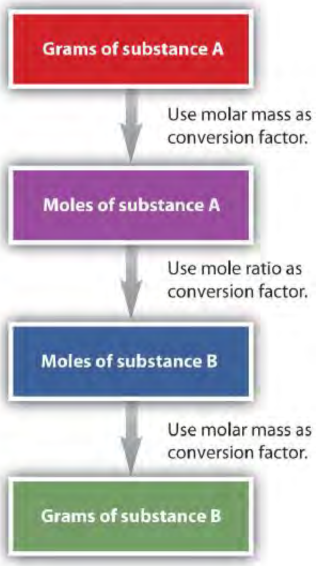
$$2\,\mathrm{AlCl_3} + 3\,\mathrm{H_2O}(\ell) \rightarrow \mathrm{Al_2O_3} + 6\,\mathrm{HCl(g)}$$

Answer

0.442 mol Al_2O_3

Mass to Mass Conversions

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques. Flowchart of mole mass calculations: To convert from grams to moles of substance A, use molar mass conversion factor; To convert from moles of substance A to moles of substance B, use the mole ratio conversion factor, and to convert from moles to grams of substance B, use molar mass conversion factor

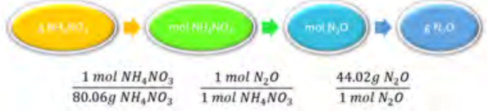
✓ Example 7.8.2: Decomposition of Ammonium Nitrate

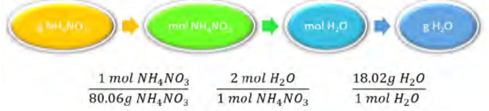
Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$\mathrm{NH_4NO_3}\,(s) \rightarrow \mathrm{N_2O}\,(g) + 2\mathrm{H_2O}\,(l)$$

In a certain experiment, $45.7\ \mathrm{g}$ of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Solutions to Example 8.5.2

<table>
<tr><th>Steps for Problem Solving</th><th>Example $7.8.2$</th></tr>
<tr><td>Identify the "given" information and what the problem is asking you to "find."</td><td>Given: $45.7\ \mathrm{g\ NH_4NO_3}$
Find:
Mass $\mathrm{N_2O} = ?\ \mathrm{g}$
Mass $\mathrm{H_2O} = ?\ \mathrm{g}$</td></tr>
<tr><td>List other known quantities.</td><td>$1\ \mathrm{mol\ NH_4NO_3} = 80.06\ \mathrm{g}$
$1\ \mathrm{mol\ N_2O} = 44.02\ \mathrm{g}$
$1\ \mathrm{mol\ H_2O} = 18.02\ \mathrm{g}$
1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O</td></tr>
<tr><td>Prepare two concept maps and use the proper conversion factor.</td><td>

Flowchart of conversion factors: 1 mole NH4NO3 to 80.06 grams NH4NO3, 1 mole N2O to 1 mole NH4NO3, 44.02 grams N2O to 1 mole N2O


Flowchart of conversion factors: 1 mole NH4NO3 to 80.06 grams NH4NO3, 2 moles H2O to 1 mole NH4NO3, 18.02 grams H2O to 1 mole H2O</td></tr>
<tr><td>Cancel units and calculate.</td><td>$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{1\ \mathrm{mol\ N_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{44.02\ \mathrm{g\ N_2O}}{1\ \mathrm{mol\ N_2O}} = 25.1$
$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{2\ \mathrm{mol\ H_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 20.6$</td></tr>
</table>

Steps for Problem Solving	Example $7.8.2$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

? Exercise $7.8.2$: Carbon Tetrachloride

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows:

$$CH_4(g) + 4\,Cl_2(g) \rightarrow CCl_2(l) + 4\,HCl(l)$$

How many grams of HCl are produced by the reaction of 100.0 g of CH_4?

Answer

908.7g HCl

Summary

- Calculations involving conversions between moles of a substance and the mass of that substance can be done using conversion factors.
- A balanced chemical reaction can be used to determine molar and mass relationships between substances.

7.9: Limiting Reactant and Theoretical Yield

Learning Objectives

- Identify the limiting reactant (limiting reagent) in a given chemical reaction.
- Calculate how much product will be produced from the limiting reactant.
- Calculate how much reactant(s) remains when the reaction is complete.

In all examples discussed thus far, the reactants were assumed to be present in stoichiometric quantities. Consequently, none of the reactants were left over at the end of the reaction. This is often desirable—as in the case of a space shuttle—where excess oxygen or hydrogen is not only extra freight to be hauled into orbit, but also an explosion hazard. More often, however, reactants are present in mole ratios that are not the same as the ratio of the coefficients in the balanced chemical equation. As a result, one or more of them will not be used up completely, but will be left over when the reaction is completed. In this situation, the amount of product that can be obtained is limited by the amount of only one of the reactants. The reactant that restricts the amount of product obtained is called the limiting reactant. The reactant that remains after a reaction has gone to completion is in excess.

Consider a nonchemical example. Assume you have invited some friends for dinner and want to bake brownies for dessert. You find two boxes of brownie mix in your pantry and see that each package requires two eggs. The balanced equation for brownie preparation is:

$$1\text{ box mix} + 2\text{ eggs} \rightarrow 1\text{ batch brownies} \tag{7.9.1}$$

If you have a dozen eggs, which ingredient will determine the number of batches of brownies that you can prepare? Because each box of brownie mix requires two eggs and you have two boxes, you need four eggs. Twelve eggs is eight more eggs than you need. Although the ratio of eggs to boxes in is 2:1, the ratio in your possession is 6:1. Hence the eggs are the ingredient (reactant) present in excess, and the brownie mix is the limiting reactant. Even if you had a refrigerator full of eggs, you could make only two batches of brownies.

Figure 7.9.1: The Concept of a Limiting Reactant in the Preparation of Brownies. For a chemist, the **balanced** chemical equation is the recipe that must be followed. 2 boxes of brownie mix and 12 eggs results in 2 batches of brownies and 8 eggs; in this case the 8 eggs are reactant present in excess

PhET Simulation: Reactants, Products and Leftovers

View this interactive simulation illustrating the concepts of limiting and excess reactants.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$H_2 + Cl_2(g) \rightarrow 2\,HCl(g)$$

The balanced equation shows that hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen non-reacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example, in the previous paragraph, complete reaction of the hydrogen would yield:

$$\text{mol HCl produced} = 3 \text{ mol } \mathrm{H_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol } \mathrm{H_2}} = 6 \text{ mol HCl}$$

Complete reaction of the provided chlorine would produce:

$$\text{mol HCl produced} = 2 \text{ mol } \mathrm{Cl_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol } \mathrm{Cl_2}} = 4 \text{ mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure $7.9.2$).

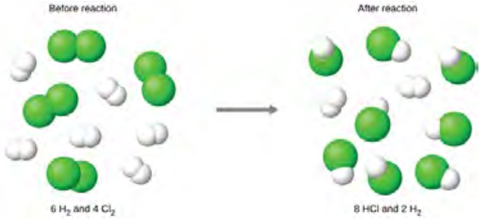


Figure $7.9.2$: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant. The figure shows a space-filling molecular models reacting. There is a reaction arrow pointing to the right in the middle. To the left of the reaction arrow there are three molecules each consisting of two green spheres bonded together. There are also five molecules each consisting of two smaller, white spheres bonded together. Above these molecules is the label, "Before reaction," and below these molecules is the label, "6 H subscript 2 and 4 C l subscript 2." To the right of the reaction arrow, there are eight molecules each consisting of one green sphere bonded to a smaller white sphere. There are also two molecules each consisting of two white spheres bonded together. Above these molecules is the label, "After reaction," and below these molecules is the label, "8 H C l and 2 H subscript 2."

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reactant; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reactant and which is in excess.

How to Identify the Limiting Reactant (Limiting Reagent)

There are two ways to determine the limiting reactant. One method is to find and compare the mole ratio of the reactants used in the reaction (Approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the **smallest** amount of product is the limiting reactant (Approach 2). This section will focus more on the second method.

Approach 1 (The "Reactant Mole Ratio Method"): Find the limiting reactant by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting (excess) reactant.

Approach 2 (The "The Product Method"): Find the limiting reactant by calculating and comparing the amount of product that each reactant will produce.

1. Balance the chemical equation for the chemical reaction.
2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reactant.
5. The reactant that produces a larger amount of product is the excess reactant.
6. To find the amount of remaining excess reactant, subtract the mass of excess reactant consumed from the total mass of excess reactant given.

The key to recognizing which reactant is the limiting reactant is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reactant. What we need to do is determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reactant. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

Example $7.9.1$: Identifying the Limiting Reactant

As an example, consider the balanced equation

$$4\,\mathrm{C_2H_3Br_3} + 11\,\mathrm{O_2} \rightarrow 8\,\mathrm{CO_2} + 6\,\mathrm{H_2O} + 6\,\mathrm{Br_2}$$

What is the limiting reactant if 76.4 grams of $C_2H_3Br_3$ reacted with 49.1 grams of O_2?

Solution

Using Approach 1:

Step 1: *Balance the chemical equation.*

The equation is already balanced with the relationship

$$4 \text{ mol } \mathrm{C_2H_3Br_3} \text{ to } 11 \text{ mol } \mathrm{O_2} \text{ to } 6 \text{ mol } \mathrm{H_2O} \text{ to } 6 \text{ mol Br}$$

Step 2: Convert all given information into moles.

$$76.4\ \text{g}\ \cancel{C_2H_3Br_3} \times \frac{1\ \text{mol}\ C_2H_3Br_3}{266.72\ \text{g}\ \cancel{C_2H_3Br_3}} = 0.286\ \text{mol}\ C_2H_3Br_3$$

$$49.1\ \text{g}\ \cancel{O_2} \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ \cancel{O_2}} = 1.53\ \text{mol}\ O_2$$

Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.

Assuming that all of the oxygen is used up,

$$1.53\ \text{mol}\cancel{O_2} \times \frac{4\ \text{mol}C_2H_3Br_3}{11\ \text{mol}\cancel{O_2}} = 0.556\ \text{mol}\ C_2H_3Br_3\ \text{are required.}$$

Because 0.556 moles of $C_2H_3Br_3$ required > 0.286 moles of $C_2H_3Br_3$ available, $\underline{C_2H_3Br_3}$ is the limiting reactant.

Using Approach 2:

Step 1: *Balance the chemical equation.*

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br_2

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$76.4\ \text{g}\ \cancel{C_2H_3Br_3} \times \frac{1\ \text{mol}\ \cancel{C_2H_3Br_3}}{266.72\ \text{g}\ \cancel{C_2H_3Br_3}} \times \frac{8\ \text{mol}\ \cancel{CO_2}}{4\ \text{mol}\ \cancel{C_2H_3Br_3}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ \cancel{CO_2}} = 25.2\ \text{g}\ CO_2$$

$$49.1\ \text{g}\ \cancel{O_2} \times \frac{1\ \text{mol}\ \cancel{O_2}}{32.00\ \text{g}\ \cancel{O_2}} \times \frac{8\ \text{mol}\ \cancel{CO_2}}{11\ \text{mol}\ \cancel{O_2}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ \cancel{CO_2}} = 49.1\ \text{g}\ CO_2$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Therefore, by either method, $C_2H_3Br_3$ is the limiting reactant.

✓ Example 7.9.2: Identifying the Limiting Reactant and the Mass of Excess Reactant

For example, in the reaction of magnesium metal and oxygen, calculate the mass of magnesium oxide that can be produced if 2.40 g Mg reacts with 10.0 g O_2. Also determine the amount of excess reactant. MgO is the only product in the reaction.

Solution

Following Approach 1:

Step 1: Balance the chemical equation.

$\underline{2}$ Mg (s) + O_2 (g) → $\underline{2}$ MgO (s)

The balanced equation provides the relationship of 2 mol Mg to 1 mol O_2 to 2 mol MgO

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$2.40\ \text{g}\ \cancel{Mg} \times \frac{1\ \text{mol}\ \cancel{Mg}}{24.31\ \text{g}\ \cancel{Mg}} \times \frac{2\ \text{mol}\ \cancel{MgO}}{2\ \text{mol}\ \cancel{Mg}} \times \frac{40.31\ \text{g}\ MgO}{1\ \text{mol}\ \cancel{MgO}} = 3.98\ \text{g}\ MgO$$

$$10.0\ \text{g}\ \cancel{O_2} \times \frac{1\ \text{mol}\ \cancel{O_2}}{32.00\ \text{g}\ \cancel{O_2}} \times \frac{2\ \text{mol}\ \cancel{MgO}}{1\ \text{mol}\ \cancel{O_2}} \times \frac{40.31\ \text{g}\ MgO}{1\ \text{mol}\ \cancel{MgO}} = 25.2\ \text{g}\ MgO$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Mg produces less MgO than does O_2 (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reactant in this reaction.

Step 5: The reactant that produces a larger amount of product is the excess reactant.

O_2 produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore O_2 is the excess reactant in this reaction.

Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reactant consumed from the total mass of excess reactant given.

Mass of excess reactant calculated using the limiting reactant:

$$2.40\ \text{g}\ \cancel{Mg} \times \frac{1\ \text{mol}\ \cancel{Mg}}{24.31\ \text{g}\ \cancel{Mg}} \times \frac{1\ \text{mol}\ \cancel{O_2}}{2\ \text{mol}\ \cancel{Mg}} \times \frac{32.00\ \text{g}\ O_2}{1\ \text{mol}\ \cancel{O_2}} = 1.58\ \text{g}\ O_2$$

OR

Mass of excess reactant calculated using the mass of the product:

$$3.98\ \text{g}\ \cancel{\text{MgO}} \times \frac{1\ \text{mol}\ \cancel{\text{MgO}}}{40.31\ \text{g}\ \cancel{\text{MgO}}} \times \frac{1\ \text{mol}\ \cancel{\text{O}_2}}{2\ \text{mol}\ \cancel{\text{MgO}}} \times \frac{32.0\ \text{g}\ \text{O}_2}{1\ \text{mol}\ \cancel{\text{O}_2}} = 1.58\ \text{g}\ \text{O}_2$$

Mass of total excess reactant given – mass of excess reactant consumed in the reaction:

10.0g O_2 - (available) 1.58g O_2 (used) = 8.42g O_2 (excess)

Therefore, O_2 is in excess.

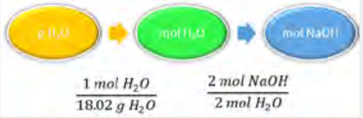
✓ Example 7.9.3: Limiting Reactant

What is the limiting reactant if 78.0 grams of Na_2O_2 were reacted with 29.4 grams of H_2O? The unbalanced chemical equation is

$$\mathrm{Na_2O_2(s) + H_2O(l) \rightarrow NaOH(aq) + H_2O_2(l)}$$

Solution

Solutions to Example 8.4.3

Steps for Problem Solving- The Product Method	Example 7.9.1
Identify the "given" information and what the problem is asking you to "find."	Given: 78.0 grams of Na_2O_2 29.4 g H_2O Find: limiting reactant
List other known quantities.	1 mol Na_2O_2= 77.96 g/mol 1 mol H_2O = 18.02 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.
Balance the equation.	Na_2O_2 (s) + $\underline{2}H_2O$ (l) → $\underline{2}$NaOH (aq) + H_2O_2 (l) The balanced equation provides the relationship of 1 mol Na_2O_2 to 2 mol H_2O 2mol NaOH to 1 mol H_2O_2
Prepare a concept map and use the proper conversion factor.	  Because the question only asks for the limiting reactant, we can perform two **mass-mole** calculations and determine which amount is less.
Cancel units and calculate.	$78.0\ \text{g Na}_2\text{O}_2 \times \frac{1\ \text{mol Na}_2\text{O}_2}{77.96\ \text{g Na}_2\text{O}_2} \times \frac{2\ \text{mol NaOH}}{1\ \text{mol Na}_2\text{O}_2} \times \frac{40\ \text{g NaOH}}{1\ \text{mol NaOH}} = 2.00\ \text{mol NaOH}$ $29.4\ \text{g H}_2\text{O} \times \frac{1\ \text{mol H}_2\text{O}}{18.02\ \text{g H}_2\text{O}} \times \frac{2\ \text{mol NaOH}}{2\ \text{mol Na}_2\text{O}_2} \times \frac{40\ \text{g NaOH}}{1\ \text{mol NaOH}} = 1.63\ \text{mol NaOH}$ Therefore, H_2O is the *limiting reactant*.

Think about your result.

✓ Example 7.9.4: Limiting Reactant and Mass of Excess Reactant

A 5.00 g quantity of Rb is combined with 3.44 g of $\mathrm{MgCl_2}$ according to this chemical reaction:

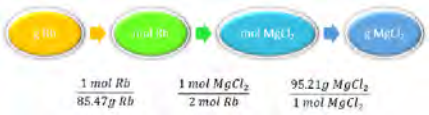
$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of Mg is formed, and what mass of remaining reactant is left over?

Solution

Solutions to Example 8.4.10

Steps for Problem Solving- The Product Method	Example 7.9.2
Steps for Problem Solving	A 5.00 g quantity of Rb is combined with 3.44 g of MgCl2 according to this chemical reaction: $2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$ What mass of Mg is formed, and what mass of remaining reactant is left over?
Identify the "given" information and what the problem is asking you to "find."	Given: 5.00g Rb, 2.44g MgCl2 Find: mass of Mg formed, mass of remaining reactant
List other known quantities.	• molar mass: Rb = 85.47 g/mol • molar mass: $MgCl_2$ = 95.21 g/mol • molar mass: Mg = 24.31 g/mol

Steps for Problem Solving- The Product Method	Example 7.9.2
Prepare concept maps and use the proper conversion factor.	Find mass Mg formed based on mass of Rb  Find mass of Mg formed based on mass of $MgCl_2$  Conversion factors: 1 mole MgCl2 to 95.21 grams MgCl2, 1 mole Mg to 1 mole MgCl2, 24.31 grams Mg to 1 mole Mg Use limiting reactant to determine amount of excess reactant consumed  Conversion factors: 1 mole Rb to 85.47 grams Rb, 1 mole MgCl2 to 2 moles Rb, 95.21 grams MgCl2 to 1 mole MgCl2
Cancel units and calculate.	Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less. $5.00\ g\,\cancel{Rb} \times \frac{1\ mol\,\cancel{Rb}}{85.47\ g\,\cancel{Rb}} \times \frac{1\ mol\,\cancel{Mg}}{2\ mol\,\cancel{Rb}} \times \frac{24.31\ g\,Mg}{1\ mol\,\cancel{Mg}} = 0.711\ g\,Mg$ $3.44\ g\,\cancel{MgCl_2} \times \frac{1\ mol\,\cancel{MgCl_2}}{95.21\ g\,\cancel{MgCl_2}} \times \frac{1\ mol\,\cancel{Mg}}{1\ mol\,\cancel{MgCl_2}} \times \frac{24.31\ g\,Mg}{1\ mol\,\cancel{Mg}} = 0.878\ g\,Mg$ The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reactant. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of $MgCl_2$ reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount. $5.00\ g\,\cancel{Rb} \times \frac{1\ mol\,\cancel{Rb}}{85.47\ g\,\cancel{Rb}} \times \frac{1\ mol\,\cancel{MgCl_2}}{2\ mol\,\cancel{Rb}} \times \frac{95.21\ g\,MgCl_2}{1\ mol\,\cancel{MgCl_2}} = 2.78\ g\,MgCl_2\ \ reacted$ Because we started with 3.44 g of $MgCl_2$, we have 3.44 g $MgCl_2$ − 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ left
Think about your result.	It usually is not possible to determine the limiting reactant using just the initial masses, as the reagents have different molar masses and coefficients.

? Exercise 7.9.1

Given the initial amounts listed, what is the limiting reactant, and what is the mass of the leftover reactant?

$$\underbrace{22.7\,g}_{MgO(s)} + \underbrace{17.9\,g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer

H_2S is the limiting reagent; 1.5 g of MgO are left over.

7.10: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants

Learning Objectives

- Calculate percentage or actual yields from known amounts of reactants.

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

Percent Yield

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, the maximum amount of product that can be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

Typically, percent yields are understandably less than 100% because of the reasons indicated earlier. However, percent yields greater than 100% are possible if the measured product of the reaction contains impurities that cause its mass to be greater than it actually would be if the product was pure. When a chemist synthesizes a desired chemical, he or she is always careful to purify the products of the reaction. Example $7.10.1$ illustrates the steps for determining percent yield.

✓ Example 7.10.1: Decomposition of Potassium Chlorate

Potassium chlorate decomposes upon slight heating in the presence of a catalyst, according to the reaction below:

$$2\text{KClO}_3\,(s) \rightarrow 2\text{KCl}\,(s) + 3\text{O}_2\,(g)$$

In a certain experiment, 40.0 g KClO_3 is heated until it completely decomposes. The experiment is performed and the oxygen gas is collected and its mass is found to be 14.9 g

a. What is the theoretical yield of oxygen gas?
b. What is the percent yield for the reaction?

Solution

a. Calculation of theoretical yield

First, we will calculate the theoretical yield based on the stoichiometry.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Mass of $\text{KClO}_3 = 40.0 \text{ g}$

Mass of O_2 collected = 14.9g

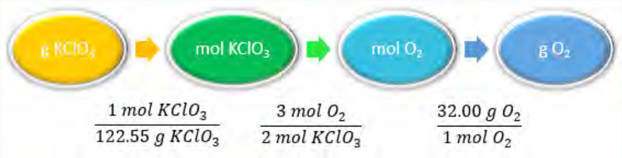
Find: Theoretical yield, g O_2

Step 2: List other known quantities and plan the problem.

1 mol $KClO_3$ = 122.55 g/mol

1 mol O_2 = 32.00 g/mol

Step 3: Apply stoichiometry to convert from the mass of a reactant to the mass of a product:



Step 4: Solve.

$$40.0\ \ \text{g}\ \cancel{\text{KClO}_3} \times \frac{1\ \ \text{mol}\ \cancel{\text{KClO}_3}}{122.55\ \ \text{g}\ \cancel{\text{KClO}_3}} \times \frac{3\ \ \text{mol}\ \cancel{\text{O}_2}}{2\ \ \text{mol}\ \cancel{\text{KClO}_3}} \times \frac{32.00\ \text{g}\ \text{O}_2}{1\ \ \text{mol}\ \cancel{\text{O}_2}} = 15.7\ \text{g}\ \text{O}_2$$

The theoretical yield of O_2 is $15.7\ \text{g}$, $15.\underline{6}7$ g unrounded.

Step 5: Think about your result.

The mass of oxygen gas must be less than the $40.0\ \text{g}$ of potassium chlorate that was decomposed.

b. Calculation of percent yield

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Theoretical yield =$15.\underline{6}7$ g, use the un-rounded number for the calculation.

Actual yield = 14.9g

Find: Percent yield, % Yield

Step 2: List other known quantities and plan the problem.

No other quantities needed.

Step 3: Use the percent yield equation below.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Step 4: Solve.

$$\text{Percent Yield} = \frac{14.9\ \text{g}}{15.\underline{6}7\ \text{g}} \times 100\% = 94.9\%$$

Step 5: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

✓ Example 7.10.2: Oxidation of Zinc

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\text{CuSO}_4(aq) + \text{Zn}(s) \rightarrow \text{Cu}(s) + \text{ZnSO}_4(aq)$$

What is the percent yield?

Solution

Solutions to Example 8.6.2

Steps for Problem Solving-The Product Method	Example $7.10.1$
Identify the "given" information and what the problem is asking you to "find."	Given: 1.274 g CuSO4 Actual yield = 0.392 g Cu Find: Percent yield
List other known quantities.	1 mol CuSO4= 159.62 g/mol 1 mol Cu = 63.55 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.
Balance the equation.	The chemical equation is already balanced. The balanced equation provides the relationship of 1 mol CuSO4 to 1 mol Zn to 1 mol Cu to 1 mol ZnSO4.
Prepare a concept map and use the proper conversion factor.	 The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield (g Cu) is found by performing **mass-mass** calculation based on the initial amount of CuSO_4.
Cancel units and calculate.	$1.274\ \text{g}\ \cancel{\text{Cu}_\text{S}\text{O}_4} \times \frac{1\ \text{mol}\ \cancel{\text{CuSO}_4}}{159.62\ \text{g}\ \cancel{\text{CuSO}_4}} \times \frac{1\ \text{mol}\ \cancel{\text{Cu}}}{1\ \text{mol}\ \cancel{\text{CuSO}_4}} \times \frac{63.55\ \text{g Cu}}{1\ \text{mol}\ \cancel{\text{Cu}}} = 0.507$ Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be: $\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$ $= \left(\frac{0.392\, g\, \text{Cu}}{0.5072\, g\, \text{Cu}}\right) \times 100$ $= 77.3\%$
Think about your result.	Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

? Exercise 7.10.1

What is the percent yield of a reaction that produces 12.5 g of the Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$\text{CCl}_4 + 2\,\text{HF} \rightarrow \text{CF}_2\text{Cl}_2 + 2\,\text{HCl}$$

Answer

48.3%

Summary

Theoretical yield is calculated based on the stoichiometry of the chemical equation. The actual yield is experimentally determined. The percent yield is determined by calculating the ratio of actual yield to theoretical yield.

7.11: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe, called the system and the surroundings. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** are everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings is the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

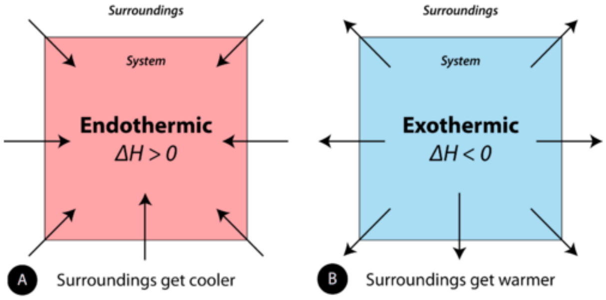


Figure 7.11.1: (A) Endothermic reaction. (B) Exothermic reaction.

Enthalpy

Heat changes in chemical reactions are often measured in the laboratory under conditions in which the reacting system is open to the atmosphere. In that case, the system is at a constant pressure. **Enthalpy** $(\boldsymbol{H})$ is the heat content of a system at constant pressure. Chemists routinely measure changes in enthalpy of chemical systems as reactants are converted into products. The heat that is absorbed or released by a reaction at constant pressure is the same as the enthalpy change, and is given the symbol ΔH. Unless otherwise specified, all reactions in this material are assumed to take place at constant pressure.

The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products. The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds. Energy needs to be put into the system in order to break chemical bonds, as they do not come apart spontaneously in most cases. Bond formation to produce products will involve release of energy. The change in enthalpy shows the trade-offs made in these two processes. Does it take more energy to break bonds than that needed to form bonds? If so, the reaction is endothermic and the enthalpy change is positive. If more energy is produced in bond formation than that needed for bond breaking, the reaction is exothermic and the enthalpy is negative.

Several factors influence the enthalpy of a system. Enthalpy is an extensive property, determined in part by the amount of material we work with. The state of reactants and products (solid, liquid, or gas) influences the enthalpy value for a system. The direction of the reaction affects the enthalpy value. A reaction that takes place in the opposite direction has the same numerical enthalpy value, but the opposite sign.

Thermochemical Equation

When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of 1 mol of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of the balanced equation:

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) + 890.4\text{ kJ}$$

The equation tells us that 1 mol of methane combines with 2 mol of oxygen to produce 1 mol of carbon dioxide and 2 mol of water. In the process, 890.4 kJ is released and so it is written as a product of the reaction. A **thermochemical equation** is a chemical equation that includes the enthalpy change of the reaction. The process in the above thermochemical equation can be shown visually in Figure 7.11.2

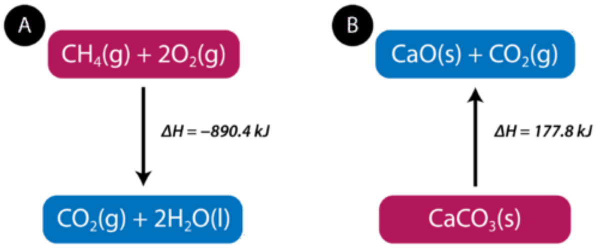


Figure 7.11.2: (A) As reactants are converted to products in an exothermic reaction, enthalpy is released into the surroundings. The enthalpy change of the reaction is negative. (B) As reactants are converted to products in an endothermic reaction, enthalpy is absorbed from the surroundings. The enthalpy change of the reaction is positive.

In the combustion of methane example, the enthalpy change is negative because heat is being released by the system. Therefore, the overall enthalpy of the system decreases. The **heat of reaction** is the enthalpy change for a chemical reaction. In the case above, the heat of reaction is -890.4 kJ. The thermochemical reaction can also be written in this way:

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) \qquad \Delta H = -890.4\ \mathrm{kJ}$$

Heats of reaction are typically measured in kilojoules. It is important to include the physical states of the reactants and products in a thermochemical equation as the value of the ΔH depends on those states.

Endothermic reactions absorb energy from the surroundings as the reaction occurs. When 1 mol of calcium carbonate decomposes into 1 mol of calcium oxide and 1 mol of carbon dioxide, 177.8 kJ of heat is absorbed. The process is shown visually in Figure 7.11.2B. The thermochemical reaction is shown below.

$$\mathrm{CaCO_3}\,(s) + 177.8\ \mathrm{kJ} \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g)$$

Because the heat is absorbed by the system, the 177.8 kJ is written as a reactant. The heat of reaction is positive for an endothermic reaction.

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \qquad \Delta H = 177.8\ \mathrm{kJ}$$

The way in which a reaction is written influences the value of the enthalpy change for the reaction. Many reactions are reversible, meaning that the product(s) of the reaction are capable of combining and reforming the reactant(s). If a reaction is written in the reverse direction, the sign of the ΔH changes. For example, we can write an equation for the reaction of calcium oxide with carbon dioxide to form calcium carbonate.

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) + 177.8\ \mathrm{kJ}$$

The reaction is exothermic and thus the sign of the enthalpy change is negative.

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) \qquad \Delta H = -177.8\ \mathrm{kJ}$$

Stoichiometric Calculations and Enthalpy Changes

Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems. Refer again to the combustion reaction of methane. Since the reaction of 1 mol of methane released 890.4 kJ, the reaction of 2 mol of methane would release $2 \times 890.4\ \mathrm{kJ} = 1781\ \mathrm{kJ}$. The reaction of 0.5 mol of methane would release $\frac{890,4\ \mathrm{kJ}}{2} = 445.2\ \mathrm{kJ}$. As with other stoichiometry problems, the moles of a reactant or product can be linked to mass or volume.

✓ Example 7.11.1

Sulfur dioxide gas reacts with oxygen to form sulfur trioxide in an exothermic reaction, according to the following thermochemical equation.

$$2\mathrm{SO_2}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{SO_3}\,(g) + 198\ \mathrm{kJ}$$

Calculate the enthalpy change that occurs when $58.0\ \mathrm{g}$ of sulfur dioxide is reacted with excess oxygen.

Solution

Step 1: List the known quantities and plan the problem.

- Mass $\mathrm{SO_2} = 58.0\ \mathrm{g}$
- Molar mass $\mathrm{SO_2} = 64.07\ \mathrm{g/mol}$
- $\Delta H = -198\ \mathrm{kJ}$ for the reaction of $2\ \mathrm{mol}\ \mathrm{SO_2}$

Unknown

- $\Delta H = ?\ \mathrm{kJ}$

The calculation requires two steps. The mass of $\mathrm{SO_2}$ is converted to moles. Then the moles of $\mathrm{SO_2}$ is multiplied by the conversion factor of $\left(\frac{-198\ \mathrm{kJ}}{2\ \mathrm{mol}\ \mathrm{SO_2}}\right)$.

Step 2: Solve.

$$\Delta H = 58.0\ \mathrm{g}\ \mathrm{SO_2} \times \frac{1\ \mathrm{mol}\ \mathrm{SO_2}}{64.07\ \mathrm{g}\ \mathrm{SO_2}} \times \frac{-198\ \mathrm{kJ}}{2\ \mathrm{mol}\ \mathrm{SO_2}} = 89.6\ \mathrm{kJ}$$

Step 3: Think about your result.

The mass of sulfur dioxide is slightly less than $1\ \mathrm{mol}$. Since $198\ \mathrm{kJ}$ is released for every $2\ \mathrm{mol}$ of $\mathrm{SO_2}$ that reacts, the heat released when about $1\ \mathrm{mol}$ reacts is one half of 198. The $89.6\ \mathrm{kJ}$ is slightly less than half of 198. The sign of ΔH is negative because the reaction is exothermic.

CHAPTER OVERVIEW

8: Gases

Of the three basic phases of matter—solids, liquids, and gases—only one of them has predictable physical properties: gases. In fact, the study of the properties of gases was the beginning of the development of modern chemistry from its alchemical roots. The interesting thing about some of these properties is that they are independent of the identity of the gas. That is, it doesn't matter if the gas is helium gas, oxygen gas, or sulfur vapors; some of their behavior is predictable and, as we will find, very similar. In this chapter, we will review some of the common behaviors of gases. Gases have no definite shape or volume; they tend to fill whatever container they are in. They can compress and expand, sometimes to a great extent. Gases have extremely low densities, one-thousandth or less the density of a liquid or solid. Combinations of gases tend to mix together spontaneously; that is, they form solutions. Air, for example, is a solution of mostly nitrogen and oxygen. Any understanding of the properties of gases must be able to explain these characteristics.

8.1: Extra-Long Straws

A drinking straw is a tube for transferring a beverage from its container to the mouth of the drinker and is typically a thin tube of plastic (such as polypropylene and polystyrene) or other material. Many people believe that when they drink a liquid they are sucking the liquid up, however the liquid is really being pushed up. A straw works because when you suck the air out of the straw, it creates a vacuum. This causes a decrease in air pressure on the inside of the straw. Since the atmospheric pressure is greater on the outside of the straw, liquid is forced into and up the straw and into your mouth (Figure 8.1.1).

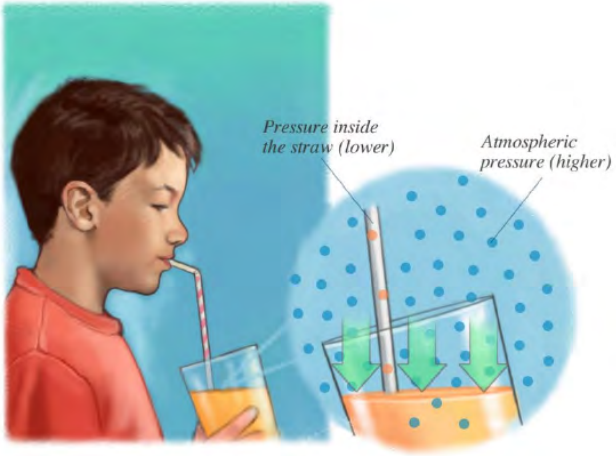


Figure 8.1.1: As you suck the air out of the straw, it creates a low pressure zone inside of it. With a low pressure zone, there is nothing pushing down on the juice, so it moves upward easily. (CC BY-NC 4.0; Ümit Kaya)

How Long of a Straw is Possible?

With the straw just sitting in the glass, the pressure on the surface of the tea is the same all over, including on the little bit of surface inside the straw. When you suck the air out of the straw, you decrease the pressure inside the straw, allowing the higher pressure on the rest of the surface to push the tea up the straw and into your mouth. Because it is really the atmosphere that is doing the pushing, the atmospheric pressure limits how high water will go up a straw.

If you formed a perfect vacuum within the straw, the pressure outside of the straw at sea level would be enough to push water to a total height of about 10.3 m. A 10.3-m column of water exerts the same pressure—101,325 N/m^2 or 14.7 lb/in^2 (psi)—as do the gas molecules in our atmosphere. At sea level, the air pressure is enough to support a column of water about thirty feet high. This means that even if you could suck all the air out of a forty-foot straw, the water would not rise more than thirty feet.

Contributions & Attributions

- Lisa Peck's Conceptual Physics Class

8.2: Kinetic Molecular Theory- A Model for Gases

Learning Objectives

- State the major concepts behind the kinetic theory of gases.
- Relate the general properties of gases to the kinetic theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas.

Figure $8.2.1$ shows a representation of how we mentally picture the gas phase.

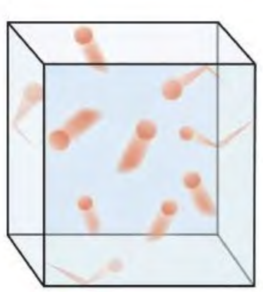


Figure $8.2.1$: The Kinetic Theory of Gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

- The physical behavior of gases is explained by the kinetic theory of gases.
- An ideal gas adheres exactly to the kinetic theory of gases.

8.3: Pressure- The Result of Constant Molecular Collisions

Learning Objectives

- Define *pressure*.
- Learn the units of pressure and how to convert between them.

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being “held in” by the earth’s gravity, rather than the gas being in a container. The pressure of the atmosphere is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

However, “average atmospheric pressure at sea level” is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mmHg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalents:

$$1\text{ atm}=760\text{ mmHg}=760\text{ torr}$$

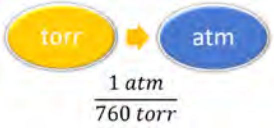
We can use these equivalents as with any equivalence—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, 1 atm = 101,325 Pa.

Example 8.3.1: Pressure Conversion

How many atmospheres are there in 595 torr?

Solution

Solutions to Example 11.3.1

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 595 torr Find: ? atm
List other known quantities.	1 atm = 760 torr
Prepare a concept map.	
Cancel units and calculate.	$595\ \cancel{torr} \times \frac{1\ atm}{760\ \cancel{torr}} = 0.783\ atm$
Think about your result.	595 torr is less than 760 torr so the final answer should be less than 1 atm.

? Exercise 8.3.1

How many atmospheres are there in 1,022 torr?

Answer

1.345 atm

Example 8.3.2: Mars

The atmosphere on Mars is largely CO_2 at a pressure of 6.01 mmHg. What is this pressure in atmospheres?

Solution

Solutions to Example 11.3.2

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 6.01mmHg Find: ? atm
List other known quantities.	1 atm = 760 mmHg
Prepare a concept map.	
Cancel units and calculate.	$6.01\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}} = 0.00791\, atm = 7.91 \times 10^{-3} atm$
Think about your result.	6.01 is a very small number relative to 760 mmHg, just like the value in atmospheres.

? Exercise 8.3.2

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer

652 torr

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

8.4: Boyle’s Law - Pressure and Volume

Learning Objectives

- Learn what is meant by the term *gas laws*.
- Learn and apply Boyle’s Law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Take pressure (*P*) and volume (*V*), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [*n*]), if the temperature (*T*) of the gas was kept constant, pressure and volume were related: as one increases, the other decreases. As one decreases, the other increases. This means that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{ constant at constant n and T}$$

If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled P_1 and V_1 and the new conditions are labeled P_2 and V_2, we have

$$P_1V_1 = \text{constant} = P_2V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1V_1 = P_2V_2 \text{ at constant n and T}$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law is called **Boyle's Law**, after the English scientist Robert Boyle, who first announced it in 1662. Figure $8.4.1$ shows two representations of how Boyle’s Law works.

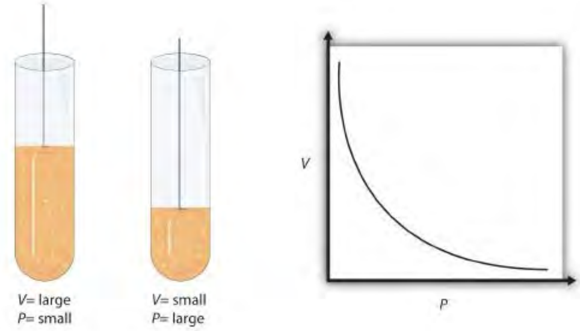


Figure $8.4.1$: Boyle’s Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot *P* versus *V* for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

Boyle’s Law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle’s Law there are two pressure variables; they must have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won’t matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

✓ Example 8.4.1

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

Solution

Solutions to Example 11.8.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 2.44 atm and V_1 = 4.01 L P_2 = 1.93 atm Find: V_2 = ? L
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{P_1 \times V_1}{P_2}$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.44 \ \cancel{\text{atm}} \times 4.01 \text{ L}}{1.93 \ \cancel{atm}} = 5.07 \text{ L}$$
Think about your result.	We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answer makes sense based on Boyle's Law.

? Exercise 8.4.1

If P_1 = 334 torr, V_1 = 37.8 mL, and P_2 = 102 torr, what is V_2?

Answer

124 mL

As mentioned, you can use any units for pressure and volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

✓ Example 8.4.2:

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

Solution

Solutions to Example 11.8.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 722 torr and V_1 = 88.8 mL V_2 = 0.633 L Find: P_2 = ? torr
List other known quantities.	1 L = 1000 mL to have the same units for volume.

Steps for Problem Solving	
Plan the problem.	1. Perform the conversion of the second volume unit from L to mL. 2. Rearrange the equation algebraically to solve for P_2. $P_2 = \frac{P_1 \times V_1}{V_2}$
Cancel units and calculate.	1. $0.663 \ \cancel{L} \times \frac{1000 \, ml}{1 \ \cancel{L}} = 663 \, ml$ 2. Substitute the known quantities into the equation and solve. $P_2 = \frac{722 \text{ torr} \times 88.8 \ \text{m}\cancel{\text{L}}}{663 \ \text{m}\cancel{\text{L}}} = 96.7 \text{ torr}$
Think about your result.	When the volume increased, the pressure decreased, which is as expected for Boyle’s Law.

? Exercise 8.4.2

If V_1 = 456 mL, P_1 = 308 torr, and P_2 = 1.55 atm, what is V_2?

Answer

119 mL

Summary

- The behavior of gases can be modeled with gas laws.
- Boyle’s Law relates the pressure and volume of a gas at constant temperature and amount.

8.5: Charles's Law- Volume and Temperature

Learning Objectives

- Learn and apply Charles's Law.

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stop of molecular motion.

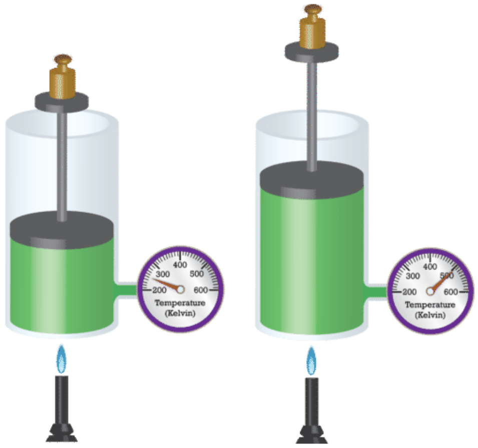


Figure 8.5.1: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 8.5.1: Temperature-Volume Data

Temperature (K)	**Volume** (mL)	$\boldsymbol{\frac{V}{T} = k} \left(\frac{\mathrm{mL}}{\mathrm{K}}\right)$
50	20	0.40
100	40	0.40
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

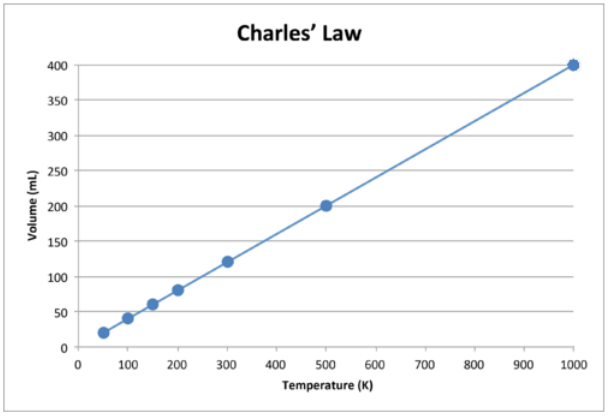


Figure 8.5.2:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\mathrm{K} = {}^\circ\mathrm{C} + 273$.

✓ Example 8.5.1:

A balloon is filled to a volume of $2.20\ \mathrm{L}$ at a temperature of $22^\mathrm{o}\mathrm{C}$. The balloon is then heated to a temperature of $71^\mathrm{o}\mathrm{C}$. Find the new volume of the balloon.

Solution

Solutions to Example 11.5.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 2.20\ \mathrm{L}$ and $T_1 = 22^\mathrm{o}\mathrm{C} = 295\ \mathrm{K}$ $T_2 = 71^\mathrm{o}\mathrm{C} = 344\ \mathrm{K}$ Find: V_2 = ? L
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times T_2}{T_1}$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.20\ \mathrm{L} \times 344\ \cancel{\mathrm{K}}}{295\ \cancel{\mathrm{K}}} = 2.57\ \mathrm{L}$$

Steps for Problem Solving	
Think about your result.	The volume increases as the temperature increases. The result has three significant figures.

Exercise 8.5.1

If V_1 = 3.77 L and T_1 = 255 K, what is V_2 if T_2 = 123 K?

Answer

1.82 L

Example 8.5.2:

A sample of a gas has an initial volume of 34.8 L and an initial temperature of −67°C. What must be the temperature of the gas for its volume to be 25.0 L?

Solution

Solutions to Example 11.5.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Given:T_1 = -27°C and V_1 = 34.8 L V_2 = 25.0 L Find: T_2 = ? K
List other known quantities.	K = -27°C + 273
Plan the problem.	1. Convert the initial temperature to Kelvin 2. Rearrange the equation algebraically to solve for T_2. $$T_2 = \frac{V_2 \times T_1}{V_1}$$
Cancel units and calculate.	1. −67°C + 273 = 206 K 2. Substitute the known quantities into the equation and solve. $$T_2 = \frac{25.0 \ \cancel{\text{L}} \times 206 \text{ K}}{34.8 \ \cancel{\text{L}}} = 148 \text{ K}$$
Think about your result.	This is also equal to −125°C. As temperature decreases, volume decreases—which it does in this example.

Exercise 8.5.2

If V_1 = 623 mL, T_1 = 255°C, and V_2 = 277 mL, what is T_2?

Answer

235 K, or −38°C

Summary

- Charles’s Law relates the volume and temperature of a gas at constant pressure and amount.

8.6: Gay-Lussac's Law- Temperature and Pressure

Learning Objectives

- Explain Gay-Lussac's Law.

Propane tanks are widely used with barbeque grills. But it's not fun to find out halfway through grilling that you have run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778-1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.

Figure 8.6.1: Joseph Gay-Lussac.

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \text{ and } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume, its pressure continually decreases until the gas condenses to a liquid.

✓ Example 8.6.1

The gas in an aerosol can is under a pressure of 3.00 atmat a temperature of 25°C. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of 845°C?

Solution

Solutions to Example 11.10.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $P_1 = 3.00 \text{ atm}$ $T_1 = 25^\circ\text{C} = 298 \text{ K}$ $T_2 = 845^\circ\text{C} = 1118 \text{ K}$ Find: $P_2 = ? \text{ atm}$
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for P_2. $P_2 = \frac{P_1 \times T_2}{T_1}$

Steps for Problem Solving	
Calculate.	Now substitute the known quantities into the equation and solve. $$P_2 = \frac{3.00 \text{ atm} \times 1118 \text{ } \cancel{\text{K}}}{298 \text{ } \cancel{\text{K}}} = 11.3 \text{ atm}$$
Think about your result.	The pressure increases dramatically due to a large increase in temperature.

Summary

- Pressure and temperature at constant volume are directly proportional.

8.7: The Combined Gas Law- Pressure, Volume, and Temperature

Learning Objectives

- Learn and apply the Combined Gas Law.

One thing we notice about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that we can propose a gas law that combines pressure, volume, and temperature. This gas law is known as the **Combined Gas Law**, and its mathematical form is

$$\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}\ at\ constant\ n$$

This allows us to follow changes in all three major properties of a gas. Again, the usual warnings apply about how to solve for an unknown algebraically (isolate it on one side of the equation in the numerator), units (they must be the same for the two similar variables of each type), and units of temperature must be in Kelvin.

Example 8.7.1:

A sample of gas at an initial volume of 8.33 L, an initial pressure of 1.82 atm, and an initial temperature of 286 K simultaneously changes its temperature to 355 K and its volume to 5.72 L. What is the final pressure of the gas?

Solution

Solutions to Example 11.4.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: V_1 = 8.33 L, P_1 = 1.82 atm, and T_1 = 286 K V2 = 5.72 L and T_2 = 355 K Find: $P2$ = ? atm
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $P_2=\frac{P_1V_1T_2}{T_1V_2}$
Calculate.	Now substitute the known quantities into the equation and solve. $P_2=\frac{(1.82\,atm)(8.33\ \cancel{L})(355\ \cancel{K})}{(286\ \cancel{K})(5.72\ \cancel{L})}=3.22atm$
Think about your result.	Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas were changing.

Exercise 8.7.1

If P_1 = 662 torr, V_1 = 46.7 mL, T_1 = 266 K, P_2 = 409 torr, and T_2 = 371 K, what is V_2?

Answer

105 mL

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be in the numerator and all by itself on one side of the equation.

Summary

- The Combined Gas Law relates pressure, volume, and temperature of a gas.

8.8: Avogadro’s Law- Volume and Moles

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. How much air should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles (or number of particles) of gas when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

where n is the number of moles of gas and k is a constant. Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

✓ Example 8.8.1

A balloon has been filled to a volume of $1.90\ \text{L}$with $0.0920\ \text{mol}$of helium gas. If $0.0210\ \text{mol}$of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

Solution

Solutions to Example 11.11.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 1.90\ \text{L}$ $n_1 = 0.0920\ \text{mol}$ Find: $V_2 = ?\ \text{L}$
List other known quantities.	Note that the final number of moles has to be calculated by adding the original number of moles to the moles of added helium. $n_2 = 0.0920 + 0.0210 = 0.1130\ \text{mol}$
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times n_2}{n_1}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{1.90\ \text{L} \times 0.1130\ \cancel{\text{mol}}}{0.0920\ \cancel{\text{mol}}} = 2.33\ \text{L}$$
Think about your result.	Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly.

? Exercise 8.8.1

A 12.8 L volume of gas contains .000498 moles of oxygen gas. At constant temperature and pressure, what volume does .0000136 moles of the gas fill?

Answer

0.350 L

Summary

- Calculations for relationships between volume and number of moles of a gas can be performed using Avogadro's Law.

8.9: The Ideal Gas Law- Pressure, Volume, Temperature, and Moles

Learning Objectives

- Explain the Ideal Gas Law.

There are a number of chemical reactions that require ammonia. In order to carry out the reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

Ideal Gas Law

The Combined Gas Law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. Putting these together leaves us with the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **Ideal Gas Law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable R for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R$$

The Ideal Gas Law is conveniently rearranged to look this way, with the multiplication signs omitted:

$$PV = nRT$$

The variable R in the equation is called the **ideal gas constant**.

Evaluating the Ideal Gas Constant

The value of R, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: kPa, atm, or mm Hg. Therefore, R can have three different values.

We will demonstrate how R is calculated when the pressure is measured in kPa. The volume of 1.00 mol of any gas at STP (Standard temperature, 273.15 K and pressure, 1 atm) is measured to be 22.414 L We can substitute $101.325 \text{ kPa}$for pressure, $22.414 \text{ L}$for volume, and $273.15 \text{ K}$for temperature into the ideal gas equation and solve for R.

$$\begin{aligned} R &= \frac{PV}{nT} \\ &= \frac{101.325 \text{ kPa} \times 22.414 \text{ L}}{1.000 \text{ mol} \times 273.15 \text{ K}} \\ &= 8.314 \text{ kPa} \cdot \text{L/K} \cdot \text{mol} \end{aligned}$$

This is the value of R that is to be used in the ideal gas equation when the pressure is given in kPa. The table below shows a summary of this and the other possible values of R. It is important to choose the correct value of R to use for a given problem.

Table 8.9.1: Values of the Ideal Gas Constant

Unit of P	Unit of V	Unit of n	Unit of T	Value and Unit of R
kPa	L	mol	K	$8.314 \text{ J/K} \cdot \text{mol}$
atm	L	mol	K	$0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}$
mm Hg	L	mol	K	$62.36 \text{ L} \cdot \text{mm Hg/K} \cdot \text{mol}$

Notice that the unit for R when the pressure is in kPa has been changed to $\text{J/K} \cdot \text{mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

Example $8.9.1$ Oxygen Gas

What volume is occupied by $3.76\ \text{g}$ of oxygen gas at a pressure of $88.4\ \text{kPa}$ and a temperature of 19^oC? Assume the oxygen is ideal.

Solution

Solutions to Example 11.5.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: • $P = 88.4\ \text{kPa}$ • $T = 19^\text{o}\text{C} = 292\ \text{K}$ Mass $\text{O}_2 = 3.76\ \text{g}$ Find: V = ? L
List other known quantities.	$\text{O}_2 = 32.00\ \text{g/mol}$ $R = 8.314\ \text{J/K} \cdot \text{mol}$
Plan the problem.	1. First, determine the number of moles of O_2 from the given mass and the molar mass. 2. Then, rearrange the equation algebraically to solve for V $V = \frac{nRT}{P}$
Calculate.	1. $3.76\ \cancel{\text{g}} \times \frac{1\ \text{mol O}_2}{32.00\ \cancel{\text{g}}\ \text{O}_2} = 0.1175\ \text{mol O}_2$ 2. Now substitute the known quantities into the equation and solve. $V = \frac{nRT}{P} = \frac{0.1175\ \cancel{\text{mol}} \times 8.314\ \text{J/}\cancel{\text{K}} \cdot \cancel{\text{mol}} \times 292\ \cancel{\text{K}}}{88.4\ \text{k}\cancel{\text{Pa}}} = 3.23\ \text{L O}_2$
Think about your result.	The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar volume $(22.4\ \text{L/mol})$ since the pressure and temperature are reasonably close to standard. The result has three significant figures because of the values for T and P. Since a joule $(\text{J}) = \text{kPa} \cdot \text{L}$, the units cancel out correctly, leaving a volume in liters.

Example $8.9.2$: Argon Gas

A 4.22 mol sample of Ar has a pressure of 1.21 atm and a temperature of 34°C. What is its volume?

Solution

Solutions to Example 11.5.2

Steps for Problem Solving

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: n = 4.22 mol P = 1.21 atm T = 34°C Find: *V* = ? L
List other known quantities.	none
Plan the problem.	1. The first step is to convert temperature to Kelvin. 2. Then, rearrange the equation algebraically to solve for V $$V = \frac{nRT}{P}$$
Calculate.	1. 34 + 273 = 307 K 2. Now substitute the known quantities into the equation and solve. $$V = \frac{(4.22\ \cancel{mol})(0.08205 \frac{L.\ \cancel{atm}}{\cancel{mol.K}})(307\ \cancel{K})}{1.21\ \cancel{atm}}$$ $$= 87.9\ L$$
Think about your result.	The number of moles of Ar is large so the expected volume should also be large.

Exercise 8.9.1

A 0.0997 mol sample of O_2 has a pressure of 0.692 atm and a temperature of 333 K. What is its volume?

Answer

3.94 L

Exercise 8.9.2

For a 0.00554 mol sample of H_2, *P* = 23.44 torr and *T* = 557 K. What is its volume?

Answer

8.21 L

Summary

- The Ideal Gas Law is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas.

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8.10: Mixtures of Gases - Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen

Learning Objectives

- Explain Dalton's Law of Partial Pressures.

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are 96.5% carbon dioxide and 3% nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus would contribute a pressure well over 2700 mm Hg. And there is no oxygen present, so we couldn't breathe there. Not that we would want to go to Venus, as the surface temperature is usually over 460^oC.

Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about 78% nitrogen and 21% oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up 78% of the gas particles in a given sample of air, it exerts 78% of the pressure. If the overall atmospheric pressure is 1.00 atm, then the pressure of just the nitrogen in the air is 0.78 atm. The pressure of the oxygen in the air is 0.21 atm.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of nitrogen is represented by P_{N_2}. **Dalton's Law of Partial Pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's Law can be expressed with the following equation:

$$P_\text{total} = P_1 + P_2 + P_3 + \cdots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, P_1 and P_2, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300 \text{ mm Hg}$ and $P_2 = 500 \text{ mm Hg}$, then $P_\text{total} = 800 \text{ mm Hg}$.

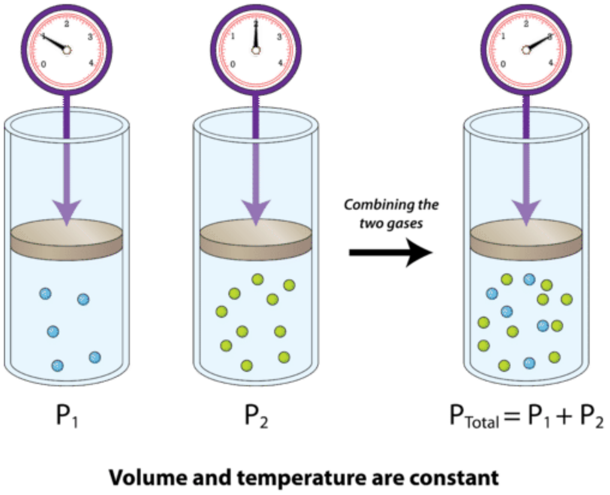

Figure 8.10.1: Dalton's Law states that the pressure of a gas mixture is equal to the partial pressures of the combining gases.

Collecting Gases Over Water

You need to do a lab experiment where hydrogen gas is generated. In order to calculate the yield of gas, you have to know the pressure inside the tube where the gas is collected. But how can you get a barometer in there? Very simple: you don't. All you need is the atmospheric pressure in the room. As the gas pushes out the water, it is pushing against the atmosphere, so the pressure inside is equal to the pressure outside.

Gas Collection by Water Displacement

Gases that are produced in laboratory experiments are often collected by a technique called **water displacement** (Figure 8.10.2). A bottle is filled with water and placed upside-down in a pan of water. The reaction flask is fitted with rubber tubing, which is then fed under the bottle of water. As the gas is produced in the reaction flask, it exits through the rubber tubing and displaces the water in the bottle. When the bottle is full of the gas, it can be sealed with a lid.

Figure 8.10.2: A gas produced in a chemical reaction can be collected by water displacement.

Because the gas is collected over water, it is not pure, but is mixed with vapor from the evaporation of the water. Dalton's Law can be used to calculate the amount of the desired gas by subtracting the contribution of the water vapor.

$$P_{\text{Total}} = P_g + P_{H_2O}$$

where P_g is the pressure of the desired gas, which can be solved for:

$$P_g = P_{Total} - P_{H_2O}$$

In order to solve a problem, it is necessary to know the vapor pressure of water at the temperature of the reaction (see table below). The sample problem illustrates the use of Dalton's Law when a gas is collected over water.

Table 8.10.1: Vapor Pressure of Water (mm Hg) at Selected Temperatures $(^\circ\text{C})$

0	**5**	**10**	**15**	**20**	**25**	**30**	**35**	**40**	**45**	**50**	**55**	**60**
4.58	6.54	9.21	12.79	17.54	23.76	31.82	42.18	55.32	71.88	92.51	118.04	149.38

✓ Example 14.14.1

A certain experiment generates 2.58 L of hydrogen gas, which is collected over water. The temperature is 20°C and the atmospheric pressure is 98.60 kPa Find the volume that the dry hydrogen would occupy at STP.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $V_{\text{Total}} = 2.58 \text{ L}$
- $T = 20^\text{o}\text{C} = 293 \text{ K}$
- $P_{\text{Total}} = 98.60 \text{ kPa} = 739.7 \text{ mm Hg}$

Unknown

- V_{H_2} at STP $= ? \text{ L}$

The atmospheric pressure is converted from kPa to mm Hg in order to match units with the table. The sum of the pressures of the hydrogen and the water vapor is equal to the atmospheric pressure. The pressure of the hydrogen is found by subtraction.

Then, the volume of the gas at STP can be calculated by using the combined gas law.

Step 2: Solve.

$$\begin{aligned} P_{H_2} &= P_{\text{Total}} - P_{H_2O} \\ &= 739,7 \text{ mm Hg} - 17.54 \text{ mm Hg} \\ &= 722.2 \text{ mm Hg} \end{aligned}$$

Now the combined gas law is used, solving for V_2, the volume of hydrogen at STP.

$$\begin{aligned} V_2 &= \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1} \\ &= \frac{722.2 \text{ mm Hg} \times 2.58 \text{ L} \times 273 \text{ K}}{760 \text{ mm Hg} \times 293 \text{ K}} \\ &= 2.28 \text{ L } \mathrm{H_2} \end{aligned}$$

Step 3: Think about your result.

If the hydrogen gas were to be collected at STP and without the presence of the water vapor, its volume would be 2.28 L. This is less than the actual collected volume because some of that is water vapor. The conversion using STP is useful for stoichiometry purposes.

Summary

- Dalton's Law of Partial Pressures states that the total pressure in a system is equal to the sum of the partial pressures of the gases present.
- The vapor pressure due to water in a sample can be corrected for, in order to get the true value for the pressure of the gas.

CHAPTER OVERVIEW

9: Solutions

Solutions play a very important role in many biological, laboratory, and industrial applications of chemistry. Of particular importance are solutions involving substances dissolved in water, or aqueous solutions. Solutions represent equilibrium systems, and the lessons learned in our last unit will be of particular importance again. Quantitative measurements of solutions are another key component of this unit. Solutions can involve all physical states - gases dissolved in gases (the air around us), solids dissolved in solids (metal alloys), liquids dissolved in solids (amalgams - liquid mercury dissolved in another metal such as silver, tin or copper). In this unit we will almost exclusively be concerned with aqueous solutions - substances dissolved in water.

9.1: Prelude - Tragedy in Cameroon

Lake Nyos is a deep crater lake in the Northwest region of Cameroon, high on the flank of an inactive volcano in the Oku volcanic plain along the Cameroon line of volcanic activity. A volcanic dam impounds the lake waters. A pocket of magma lies beneath the lake and leaks carbon dioxide (CO_2) into the water, changing it into carbonic acid. Nyos is one of only three known exploding lakes to be saturated with carbon dioxide in this way.

In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (Figure 9.1.1), a deep lake in a volcanic crater. It is believed that the lake underwent a turnover due to gradual heating from below the lake, and the warmer, less-dense water saturated with carbon dioxide reached the surface. Consequently, tremendous quantities of dissolved CO_2 were released, and the colorless gas, which is denser than air, flowed down the valley below the lake and suffocated humans and animals living in the valley.

Figure 9.1.1: Two photos are shown. The first is an aerial view of a lake surrounded by green hills. The second shows a large body of water with a fountain sending liquid up into the air several yards or meters above the surface of the water. (a) It is believed that the 1986 disaster, that killed more than 1700 people near Lake Nyos in Cameroon, resulted when a large volume of carbon dioxide gas was released from the lake. (b) A CO_2 vent has since been installed to help outgas the lake in a slow, controlled fashion and prevent a similar catastrophe from happening in the future. (Credit a: modification of work by Jack Lockwood; credit b: modification of work by Bill Evans.)

Following the Lake Nyos tragedy, scientists investigated other African lakes to see if a similar phenomenon could happen elsewhere. Lake Kivu in Democratic Republic of Congo, 2,000 times larger than Lake Nyos, was also found to be supersaturated, and geologists found evidence for out-gassing events around the lake about every one thousand years.

Contributions & Attributions

- Wikipedia
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9.2: Solutions - Homogeneous Mixtures

Learning Objectives

- Learn terminology involving solutions.
- Explain the significance of the statement "like dissolves like."
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater or lesser presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious.

Figure 9.2.1: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent. Table 9.2.1 lists some common types of solutions, with examples of each.

Table 9.2.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	salt water
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

Example 9.2.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

? Exercise 9.2.1

A solution is made by dissolving 3.33 g of $\mathrm{HCl(g)}$ in 40.0 g of liquid methyl alcohol $(\mathrm{CH_3OH})$. Identify the solvent and solute in the resulting solution.

Answer

solute: $\mathrm{HCl(g)}$

solvent: $\mathrm{CH_3OH}$

Like Dissolves Like

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, does dissolve in water—it is this oxygen that the fish take in through their gills. The reason we can enjoy carbonated sodas is also due to a nonpolar compound that dissolves in water. Pepsi-cola and all the other sodas have carbon dioxide gas, $\mathrm{CO_2}$, a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Table 9.2.2: Summary of Solubilities

Solute (Polarity of Compound)	Solvent (Polarity of Compound)	Dominant Intermolecular Force	Is Solution Formed?
Polar	Polar	Dipole-Dipole Force and/or Hydrogen Bond	yes
Non-polar	Non-polar	Dispersion Force	yes
Polar	Non-polar		no
Non-polar	Polar		no
Ionic	Polar	Ion-Dipole	yes
Ionic	Non-polar		no

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges, a solution does not form. When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in Chapter 10). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water.

Figure 9.2.2: Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

✓ Example 9.2.2: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O? Explain your answer.

Solution

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

? Exercise 9.2.2

Would C_3H_7OH be more soluble in CCl_4 or H_2O? Explain your answer.

Answer

H_2O, because both experience hydrogen bonding.

✓ Example 9.2.3

Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol (CH_3OH)
b. sodium sulfate (\ce{Na2SO4}\))
c. octane (C_8H_{18})

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

a. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
b. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
c. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

? Exercise 9.2.3: Toluene

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer

> Octane (C_8H_{18}) will dissolve. It is also non-polar.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.

9.3: Solutions of Solids Dissolved in Water- How to Make Rock Candy

Learning Objectives

- Define electrolytes and non electrolytes
- Explain why solutions form.
- Discuss the idea of water as the "universal solvent".
- Explain how water molecules attract ionic solids when they dissolve in water.

We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. We also know that solutions have constant composition, and that this composition can be varied up to a point to maintain the homogeneous nature of the solution. But how exactly do solutions form? Why is it that oil and water will not form a solution, and yet vinegar and water will? Why could we dissolve table salt in water, but not in vegetable oil? The reasons why solutions will form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Solubility and Saturation

Table salt (NaCl) readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is specified as the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table $9.3.1$ lists the solubilities of some simple ionic compounds. These solubilities vary widely. NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table $9.3.1$: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated, but rather concentrated. In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

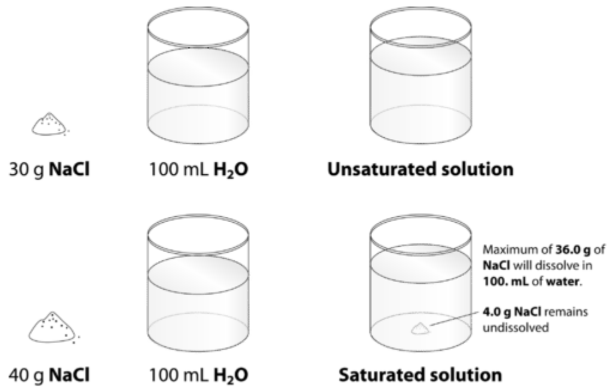


Figure 9.3.1: When 30.0 g of $NaCl$ is added to 100 mL, it all dissolves, forming an unsaturated solution. When 40.0 g is added, 36.0 g dissolves and 4.0 g remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

Electrolyte Solutions: Dissolved Ionic Solids

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte (good conductor). If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, the substance is a weak electrolyte (does not conduct electricity as well).

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 9.3.1).

Figure 9.3.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte. This diagram shows three separate beakers. Each has a wire plugged into a wall outlet. In each case, the wire leads from the wall to the beaker and is split resulting in two ends. One end leads to a light bulb and continues on to a rectangle labeled with a plus sign. The other end leads to a rectangle labeled with a minus sign. The rectangles are in a solution. In the first beaker, labeled “Ethanol No Conductivity,” four pairs of linked small green spheres suspended in the solution between the rectangles. In the second beaker, labeled “K C l Strong Conductivity,” six individual green spheres, three labeled plus and three labeled minus are suspended in the solution. Each of the six spheres has an arrow extending from it pointing to the rectangle labeled with the opposite sign. In the third beaker, labeled “Acetic acid solution Weak conductivity,” two pairs of joined green spheres and two individual spheres, one labeled plus and one labeled minus are shown suspended between the two rectangles. The plus labeled sphere has an arrow pointing to the rectangle labeled minus and the minus labeled sphere has an arrow pointing to the rectangle labeled plus.

Water and other polar molecules are attracted to ions, as shown in Figure 9.3.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

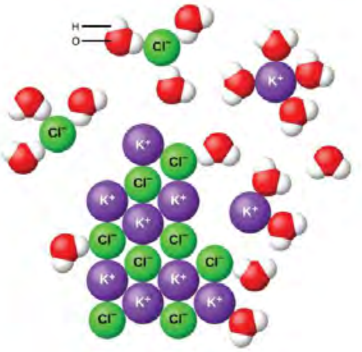


Figure 9.3.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown. The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled C l superscript minus mixed and touching near the center of the diagram. Outside of this cluster of spheres are seventeen clusters of three spheres, which include one red and two white spheres. A red sphere in one of these clusters is labeled O. A white sphere is labeled H. Two of the green C l superscript minus spheres are surrounded by three of the red and white clusters, with the red spheres closer to the green spheres than the white spheres. One of the K superscript plus purple spheres is surrounded by four of the red and white clusters. The white spheres of these clusters are closest to the purple spheres.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

✓ Example 9.3.1: Identifying Ionic Compounds

Which compound(s) will dissolve in solution to separate into ions?

a. LiF

b. P_2F_5
c. C_2H_5OH

Solution

LiF will separate into ions when dissolved in solution, because it is an ionic compound. P_2F_5 and C_2H_5OH are both covalent and will stay as molecules in a solution.

? Exercise 9.3.1

Which compounds will dissolve in solution to separate into ions?

a. $C_6H_{12}O_{11}$, glucose
b. CCl_4
c. $CaCl_2$
d. $AgNO_3$

Answer

c & d

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per 100 g of solvent. The solubility of sodium chloride in water is 36.0 g per 100 g water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (Figure 9.3.4).

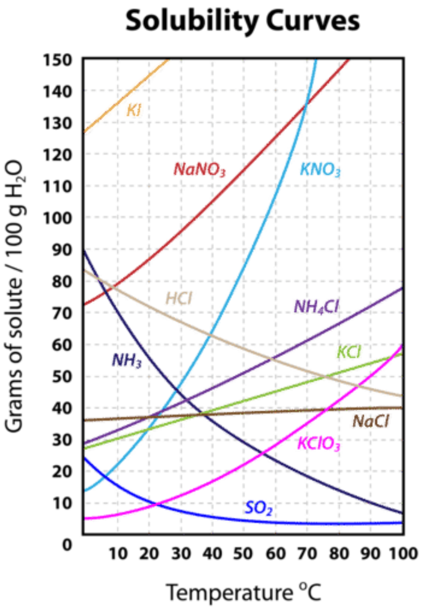


Figure 9.3.4: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, $\mathrm{NH_3}$, and $\mathrm{SO_2}$—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that 80 g of $\mathrm{KNO_3}$ is added to 100 g of water at $30^\mathrm{o}\mathrm{C}$. According to the solubility curve, approximately 48 g of $\mathrm{KNO_3}$ will dissolve at $30^\mathrm{o}\mathrm{C}$. This means that the solution will be saturated since 48 g is less than 80 g. We can also determine that there will be $80-48=32$ g of undissolved $\mathrm{KNO_3}$ remaining at the bottom of the container. Now suppose that this saturated solution is heated to $60^\mathrm{o}\mathrm{C}$. According to the curve, the solubility of $\mathrm{KNO_3}$ at $60^\mathrm{o}\mathrm{C}$ is about 107 g. Now the solution is unsaturated since it contains only the original 80 g of dissolved solute. Now suppose the solution is cooled all the way down to $0^\mathrm{o}\mathrm{C}$. The solubility at $0^\mathrm{o}\mathrm{C}$ is about 14 g, meaning that $80-14=66$ g of the $\mathrm{KNO_3}$ will re-crystallize.

Summary

- Solubility is the specific amount of solute that can dissolve in a given amount of solvent.
- Saturated and unsaturated solutions are defined.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).
- Water is considered the universal solvent since it can dissolve both ionic and polar solutes, as well as some nonpolar solutes (in very limited amounts).
- The solubility of a solid in water increases with an increase in temperature.

Vocabulary

- **Miscible** - Liquids that have the ability to dissolve in each other.
- **Immiscible** - Liquids that do not have the ability to dissolve in each other.
- **Electrostatic attraction** - The attraction of oppositely charged particles.

9.4: Solutions of Gases in Water

Learning Objectives

- Explain how temperature and pressure affect the solubility of gases.

In an earlier module of this chapter, the effect of intermolecular attractive forces on solution formation was discussed. The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility. For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium, but 100 times less than the solubility of chloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the liquid hydrocarbon hexane, C_6H_{14}, is approximately 20 times greater than it is in water.

Other factors also affect the solubility of a given substance in a given solvent. Temperature is one such factor, with gas solubility typically decreasing as temperature increases (Figure $9.4.1$). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

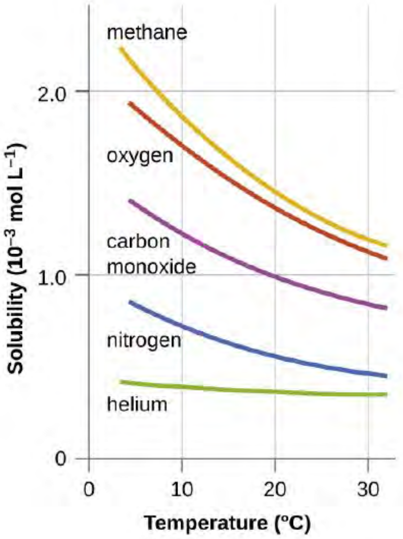


Figure $9.4.1$: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (Figure $9.4.2$).

(a) (b)

Figure $9.4.2$: (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (Credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service.)

The solubility of a gaseous solute is also affected by the partial pressure of solute in the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at this pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution in the form of small bubbles (Figure $9.4.3$). At this point, the beverage is *supersaturated* with carbon dioxide and, with time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become “flat.”

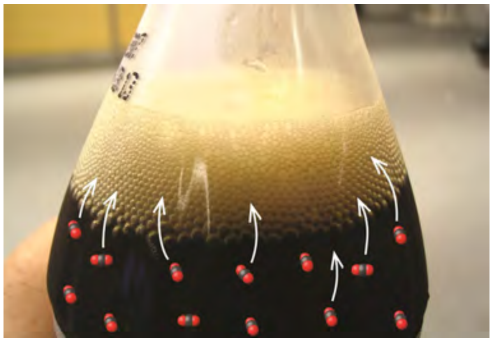


Figure $9.4.3$: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. A dark brown liquid is shown in a clear, colorless container. A thick layer of beige bubbles appear at the surface of the liquid. In the liquid, thirteen small clusters of single black spheres with two red spheres attached to the left and right are shown. Red spheres represent oxygen atoms and black represent carbon atoms. Seven white arrows point upward in the container from these clusters to the bubble layer at the top of the liquid. (Credit: modification of work by Derrick Coetzee.)

"Fizz"

The dissolution in a liquid, also known as fizz, usually involves carbon dioxide under high pressure. When the pressure is reduced, the carbon dioxide is released from the solution as small bubbles, which causes the solution to become effervescent, or fizzy. A common example is the dissolving of carbon dioxide in water, resulting in carbonated water.

Carbon dioxide is weakly soluble in water, therefore it separates into a gas when the pressure is released. This process is generally represented by the following reaction, where a pressurized dilute solution of carbonic acid in water releases gaseous carbon dioxide at decompression:

$$\mathrm{H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)}$$

In simple terms, it is the result of the chemical reaction occurring in the liquid which produces a gaseous product.

For many gaseous solutes, the relation between solubility, C_g, and partial pressure, P_g, is a proportional one:

$$C_{\mathrm{g}} = kP_{\mathrm{g}}$$

where k is a proportionality constant that depends on the identities of the gaseous solute and solvent, and on the solution temperature. This is a mathematical statement of Henry's law: *The quantity of an ideal gas that dissolves in a definite volume of liquid is directly proportional to the pressure of the gas.*

✓ Example 9.4.1: Application of Henry's Law

At 20 °C, the concentration of dissolved oxygen in water exposed to gaseous oxygen at a partial pressure of 101.3 kPa (760 torr) is 1.38 × 10^{-3} mol L^{-1}. Use Henry's law to determine the solubility of oxygen when its partial pressure is 20.7 kPa (155 torr), the approximate pressure of oxygen in earth's atmosphere.

Solution

According to Henry's law, for an ideal solution the solubility, C_g, of a gas (1.38 × 10^{-3} mol L^{-1}, in this case) is directly proportional to the pressure, P_g, of the undissolved gas above the solution (101.3 kPa, or 760 torr, in this case). Because we know both C_g and P_g, we can rearrange this expression to solve for k.

$$\begin{aligned} C_{\mathrm{g}} &= kP_{\mathrm{g}} \\ k &= \frac{C_{\mathrm{g}}}{P_{\mathrm{g}}} \\ &= \frac{1.38 \times 10^{-3}\ \mathrm{mol\ L^{-1}}}{101.3\ \mathrm{kPa}} \\ &= 1.36 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ kPa^{-1}} \\ &\quad (1.82 \times 10^{-6}\ \mathrm{mol\ L^{-1}\ torr^{-1}}) \end{aligned}$$

Now we can use k to find the solubility at the lower pressure.

$$C_{\mathrm{g}} = kP_{\mathrm{g}}$$

$$\begin{aligned} &1.36 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ kPa^{-1}} \times 20.7\ \mathrm{kPa} \\ &(\mathrm{or}\ 1.82 \times 10^{-6}\ \mathrm{mol\ L^{-1}\ torr^{-1}} \times 155\ \mathrm{torr}) \\ &= 2.82 \times 10^{-4}\ \mathrm{mol\ L^{-1}} \end{aligned}$$

Note that various units may be used to express the quantities involved in these sorts of computations. Any combination of units that yield to the constraints of dimensional analysis are acceptable.

? Exercise 9.4.1

A 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 20.26 kPa (152 torr) resulted in the dissolution of 1.45 × 10^{-3} g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 101.3 kPa (760 torr).

Answer

7.25 × 10^{-3} g

Case Study: Decompression Sickness ("The Bends")

Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing an increase of approximately 1 atm for each 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher ambient pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.

As the diver ascends to the surface of the water, the ambient pressure decreases and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that can cause a variety of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompression stops, pausing for several minutes at given depths during the ascent. When these preventative measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (Figure $9.4.4$).

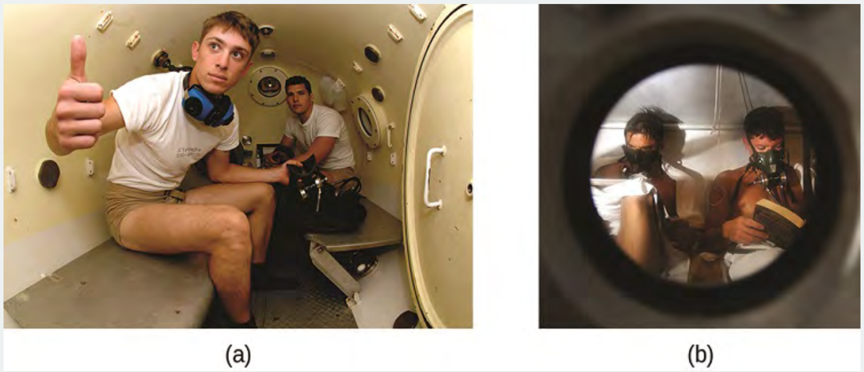
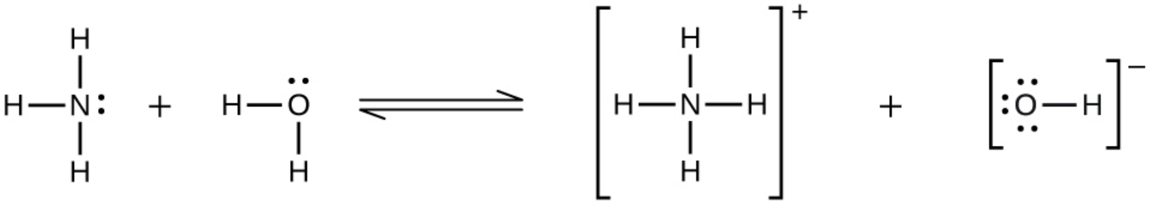


Figure $9.4.4$: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy. Two photos are shown. The first shows two people seated in a steel chamber on benches that run length of the chamber on each side. The chamber has a couple of small circular windows and an open hatch-type door. One of the two people is giving a thumbs up gesture. The second image provides a view through a small, circular window. Inside the two people can be seen with masks over their mouths and noses. The people appear to be reading.

Deviations from Henry's law are observed when a chemical reaction takes place between the gaseous solute and the solvent. Thus, for example, the solubility of ammonia in water does not increase as rapidly with increasing pressure as predicted by the law because ammonia, being a base, reacts to some extent with water to form ammonium ions and hydroxide ions.



This reaction diagram shows three H atoms bonded to an N atom above, below, and two the left of the N. A single pair of dots is present on the right side of the N. This is followed by a plus, then two H atoms bonded to an O atom to the left and below the O. Two pairs of dots are present on the O, one above and the other to the right of the O. A double arrow, with a top arrow pointing right and a bottom arrow pointing left follows. To the right of the double arrow, four H atoms are shown bonded to a central N atom. These 5 atoms are enclosed in brackets with a superscript plus outside. A plus follows, then an O atom linked by a bond to an H atom on its right. The O atom has pairs of dots above, to the left, and below the atom. The linked O and H are enclosed in brackets with superscript minus outside.

Gases can form supersaturated solutions. If a solution of a gas in a liquid is prepared either at low temperature or under pressure (or both), then as the solution warms or as the gas pressure is reduced, the solution may become supersaturated.

Contributions & Attributions

- Wikipedia
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

9.5: Solution Concentration- Mass Percent

Learning Objectives

- Express the amount of solute in a solution in various concentration units.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms with meanings that depend on various factors.

Introduction

Concentration is the measure of how much of a given substance is mixed with another substance. Solutions are said to be either dilute or concentrated. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution.

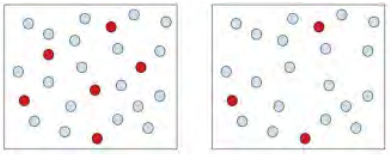


Figure $9.5.1$: The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. To make juice, you have to mix the frozen juice concentrate from inside these containers with three or four times the container size full of water. Therefore, you are diluting the concentrated juice. In terms of solute and solvent, the concentrated solution has a lot of solute versus the dilute solution that would have a smaller amount of solute.

The terms "concentrated" and "dilute" provide qualitative methods of describing concentration. Although qualitative observations are necessary and have their place in every part of science, including chemistry, we have seen throughout our study of science that there is a definite need for quantitative measurements in science. This is particularly true in solution chemistry. In this section, we will explore some quantitative methods of expressing solution concentration.

Mass Percent

There are several ways of expressing the concentration of a solution by using a percentage. The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \mathrm{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving $25.0\ \mathrm{g}$ of sugar into $100.0\ \mathrm{g}$ of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\text{ g sugar}}{125.0\text{ g solution}} \times 100\% = 20.0\%\text{ sugar}$$

Example 9.5.1

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\% \text{ m/m} = \frac{36.5 \text{ g}}{355 \text{ g}} \times 100\% = 10.3\%$$

Exercise 9.5.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer

7.90 %

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g \text{ solute}}{100 \text{ gsolution}}$ or $\frac{100 \text{ gsolution}}{g \text{ solute}}$

For example, if you need to make 3000.0 gof a 5.00% solution of sodium chloride, the mass of solute needs to be determined.

Solution

Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0 \; g \; NaCl \; \cancel{solution} \times \frac{5.00 \; g \; NaCl}{100 \; g \; NaCl \; \cancel{solution}} = 150.0g \; NaCl$$

You would need to weigh out 150 g of NaCl and add it to 2850 g of water. Notice that it was necessary to subtract the mass of the NaCl (150 g) from the mass of solution (3000 g) to calculate the mass of the water that would need to be added.

? Exercise 9.5.1

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg, 3.00 % (by mass) H_2O_2 solution?

Answer

180 g H_2O_2

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9.6: Solution Concentration- Molarity

Learning Objectives

- Use molarity to determine quantities in chemical reactions.
- Use molarity as a conversion factor in calculations.

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{9.6.1}$$

The symbol for molarity is M or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $\left[\text{Ag}^{+}\right]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to perform calculations with, but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” This is read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> *Be sure to note that molarity is calculated as the total volume of the* ***entire*** *solution, not just volume of solvent! The solute contributes to total volume.*

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

Step 1: First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ g\cancel{HCl} \times \frac{1\ mol\ HCl}{36.5\ g\cancel{HCl}} = 0.614\, mol\ HCl$$

Step 2: Now we can use the definition of molarity to determine a concentration:

$$M\ = \frac{0.614\, mol\ HCl}{1.56 L\ solution} = 0.394\, M HCl$$

Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

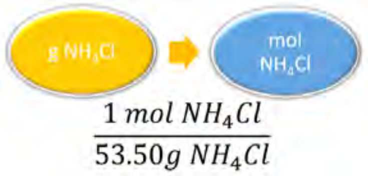
Example 9.6.1

A solution is prepared by dissolving $42.23 \text{ g}$of NH_4Cl into enough water to make $500.0 \text{ mL}$of solution. Calculate its molarity.

Solution

Solutions to Example 13.6.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Mass $= 42.23 \text{ g NH}_4\text{Cl}$ Volume solution $= 500.0 \text{ mL} = 0.5000 \text{ L}$ Find: Molarity = ? M
List other known quantities.	Molar mass $\text{NH}_4\text{Cl} = 53.50 \text{ g/mol}$

Steps for Problem Solving	
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  $\frac{1\ mol\ NH_4Cl}{53.50g\ NH_4Cl}$ 2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters. $\mathrm{M} = \frac{\mathrm{mol\ NH_4Cl}}{\mathrm{L\ solution}}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \ \cancel{\mathrm{g\ NH_4Cl}} \times \frac{1\ \mathrm{mol\ NH_4Cl}}{53.50\ \ \cancel{\mathrm{g\ NH_4Cl}}} = 0.7893\ \mathrm{mol\ NH_4Cl}$ $\frac{0.7893\ \mathrm{mol\ NH_4Cl}}{0.5000\ \mathrm{L\ solution}} = 1.579\ \mathrm{M}$
Think about your result.	The molarity is $1.579\ \mathrm{M}$, meaning that a liter of the solution would contain $1.579\ \mathrm{mol\ NH_4Cl}$. Four significant figures are appropriate.

Exercise $9.6.1A$

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer

1.57 M $C_6H_{12}O_6$

Exercise $9.6.1B$

What is the concentration, in $\mathrm{mol/L}$, where $137\ \mathrm{g}$ of NaCl has been dissolved in enough water to make $500\ \mathrm{mL}$ of solution?

Answer

4.69 M NaCl

Using Molarity in Calculations

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition.

Example $9.6.2$: Determining Moles of Solute, Given the Concentration and Volume of a Solution

For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

Solution

$$\frac{.887\ mol\ NaCl}{1\ L\ solution}$$

$$0.108\ L\,\cancel{solution} \times \frac{0.887\,mol\,NaCl}{1L\,\cancel{solution}} = 0.0958\,mol\,NaCl$$

If we used the definition approach, we get the same answer, but now we are using conversion factor skills.

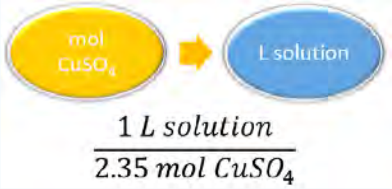
Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

✓ Example 9.6.3: Determining Volume of a Solution, Given the Concentration and Moles of Solute

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:



$$4.88\ mol\,\cancel{CuSO_4} \times \frac{1\,L\,solution}{2.35\ mol\,\cancel{CuSO_4}} = 2.08\,L\,of\ solution$$

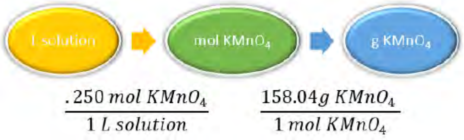
In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams. The following example illustrates this.

✓ Example 9.6.4

A chemist needs to prepare $3.00\ \text{L}$ of a $0.250\ \text{M}$ solution of potassium permanganate $\left(\text{KMnO}_4\right)$. What mass of KMnO_4 does she need to make the solution?

Solution

Solutions to Example 13.6.4

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Molarity $= 0.250\ \text{M}$ Volume $= 3.00\ \text{L}$ Find: Mass $\text{KMnO}_4 = ?\ \text{g}$
List other known quantities.	Molar mass $\text{KMnO}_4 = 158.04\ \text{g/mol}$ 0.250 mol $KMnO_4$ to 1 L of $KMnO_4$ solution
Plan the problem.	
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $\text{mol KMnO}_4 = 0.250\ \text{M KMnO}_4 \times 3.00\ \text{L} = 0.750\ \text{mol K}$ $3.00\ L\ \cancel{solution} \times \frac{0.250\ \text{mol}\ \cancel{\text{KMnO}_4}}{1\ L\ \cancel{solution}} \times \frac{158.04\ \text{g KMnO}_4}{1\ \text{mol}\ \cancel{\text{KMnO}_4}} = 119\ \text{g KMnO}$
Think about your result.	When $119\ \text{g}$ of potassium permanganate is dissolved into water to make $3.00\ \text{L}$ of solution, the molarity is $0.250\ \text{M}$.

? Exercise $9.6.4A$

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer

1.74 L

? Exercise $9.6.4B$

Answer the problems below using concentration as a conversion factor.

a. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
b. What volume of 1.50 M HCl solution contains 10.0 g of hydrogen chloride?

Answer a

6.12 g

Answer b

183 mL or 0.183L

How to Indicate Concentration

- Square brackets are often used to represent concentration, e.g., [NaOH] = 0.50 M.
- Use the capital letter M for molarity, not a lower case m (this is a different concentration unit called *molality*).

Watch as the Flinn Scientific Tech Staff demonstrates "How To Prepare Solutions."

It is important to note that there are many different ways you can set up and solve your chemistry equations. Some students prefer to answer multi-step calculations in one long step, while others prefer to work out each step individually. Neither method is necessarily better or worse than the other method—whichever makes the most sense to ***you*** is the one that you should use. In this text, we will typically use unit analysis (also called dimension analysis or factor analysis).

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

9.7: Solution Dilution

Learning Objectives

- Explain how concentrations can be changed in the lab.
- Understand how stock solutions are used in the laboratory.

We are often concerned with how much solute is dissolved in a given amount of solution. We will begin our discussion of solution concentration with two related and relative terms: **dilute** and **concentrated**.

- A **dilute** solution is one in which there is a relatively small amount of solute dissolved in the solution.
- A **concentrated** solution contains a relatively large amount of solute.

These two terms do not provide any quantitative information (actual numbers), but they are often useful in comparing solutions in a more general sense. These terms also do not tell us whether or not the solution is saturated or unsaturated, or whether the solution is "strong" or "weak". These last two terms will have special meanings when we discuss acids and bases, so be careful not to confuse them.

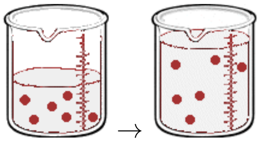
Stock Solutions

It is often necessary to have a solution with a concentration that is very precisely known. Solutions containing a precise mass of solute in a precise volume of solution are called **stock (or standard) solutions.** To prepare a standard solution, a piece of lab equipment called a volumetric flask should be used. These flasks range in size from 10 mL to 2000 mL and are carefully calibrated to a single volume. On the narrow stem is a **calibration mark**. The precise mass of solute is dissolved in a bit of the solvent, and this is added to the flask. Then, enough solvent is added to the flask until the level reaches the calibration mark.

Often, it is convenient to prepare a series of solutions of known concentrations by first preparing a single **stock solution**, as described in the previous section. **Aliquots** (carefully measured volumes) of the stock solution can then be diluted to any desired volume. In other cases, it may be inconvenient to weigh a small mass of sample accurately enough to prepare a small volume of a dilute solution. Each of these situations requires that a solution be diluted to obtain the desired concentration.

Dilutions of Stock (or Standard) Solutions

Imagine we have a salt water solution with a certain concentration. That means we have a certain amount of salt (a certain mass or a certain number of moles) dissolved in a certain volume of solution. Next, we will dilute this solution. This is done by adding more water, not more salt:



Before Dilution and After Dilution

The molarity of solution 1 is

$$M_1 = \frac{\text{moles}_1}{\text{liter}_1}$$

and the molarity of solution 2 is

$$M_2 = \frac{\text{moles}_2}{\text{liter}_2}$$

rearrange the equations to find moles:

$$\text{moles}_1 = M_1 \text{liter}_1$$

and

$$\text{moles}_2 = M_2 \text{liter}_2$$

What stayed the same and what changed between the two solutions? By adding more water, we changed the volume of the solution. Doing so also changed its concentration. **However, the number of moles of solute did not change.** So,

$$moles_1 = moles_2$$

Therefore

$$\boxed{M_1V_1 = M_2V_2} \tag{9.7.1}$$

where

- M_1 and M_2 are the concentrations of the original and diluted solutions
- V_1 and V_2 are the volumes of the two solutions

Preparing dilutions is a common activity in the chemistry lab and elsewhere. Once you understand the above relationship, the calculations are simple.

Suppose that you have $100.\ \text{mL}$ of a $2.0\ \text{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \text{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2.

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \text{M} \times 100.\ \text{mL}}{500.\ \text{mL}} = 0.40\ \text{M HCl}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves calculating what amount of a highly concentrated solution is required to make a desired quantity of solution of lesser concentration. The highly concentrated solution is typically referred to as the stock solution.

✓ Example 9.7.1: Diluting Nitric Acid

Nitric acid $\left(\text{HNO}_3\right)$ is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \text{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \text{L}$ of a $0.50\ \text{M}$ solution?

Solution

Solutions to Example13.7.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: M_1, Stock $\text{HNO}_3 = 16\ \text{M}$ $V_2 = 8.00\ \text{L}$ $M_2 = 0.50\ \text{M}$ Find: Volume stock $\text{HNO}_3\ (V_1) = ?\ \text{L}$
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_1. $$V_1 = \frac{M_2 \times V_2}{M_1}$$
Calculate and cancel units.	Now substitute the known quantities into the equation and solve. $$V_1 = \frac{0.50\ \text{M} \times 8.00\ \text{L}}{16\ \text{M}} = 0.25\ \text{L} \tag{9.7.2}$$
Think about your result.	$0.25\ \text{L}\ (250\ \text{mL})$ of the stock HNO_3 needs to be diluted with water to a final volume of $8.00\ \text{L}$. The dilution is by a factor of 32 to go from $16\ \text{M}$ to $0.5\ \text{M}$.

? Exercise 9.7.1

A 0.885 M solution of KBr with an initial volume of 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

Answer

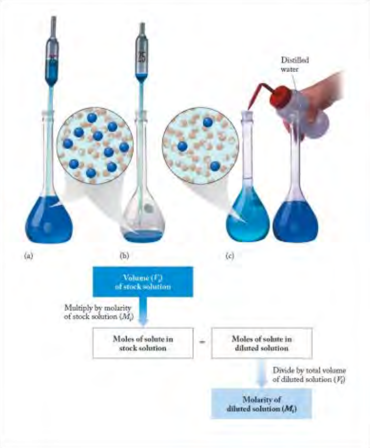
135.4 mL

Note that the calculated volume will have the same dimensions as the input volume, and dimensional analysis tells us that in this case we don't need to convert to liters, since L cancels when we divide M (mol/L) by M (mol/L).

Diluting and Mixing Solutions

Diluting and Mixing Solutions

How to Dilute a Solution by CarolinaBiological

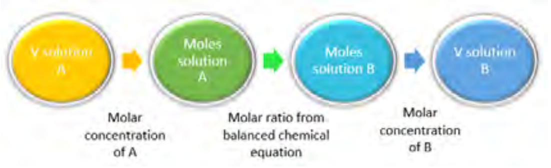


9.8: Solution Stoichiometry

Learning Objectives

- Determine amounts of reactants or products in aqueous solutions.

As we learned previously, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are “switched” (they *replace* each other). Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and therefore their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead (II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead (II) chloride.

$$\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow PbCl_2(s) + 2\,NaNO_3(aq)} \tag{9.8.1}$$

In the reaction shown above, if we mixed 0.123 L of a 1.00 M solution of NaCl with 1.50 M solution of $\mathrm{Pb(NO_3)_2}$, we could calculate the volume of $\mathrm{Pb(NO_3)_2}$ solution needed to completely precipitate the $\mathrm{Pb^{2+}}$ ions.

The molar concentration can also be expressed as the following:

$$1.00\ M\ \mathrm{NaCl} = \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}}$$

and

$$1.50\ M\,\mathrm{Pb(NO_3)_2} = \frac{1.50\ mol\ \mathrm{Pb(NO_3)_2}}{1\ L\ \mathrm{Pb(NO_3)_2 solution}}$$

First, we must examine the reaction stoichiometry in the balanced reaction (Equation 9.8.1). In this reaction, one mole of $\mathrm{Pb(NO_3)_2}$ reacts with two moles of NaCl to give one mole of $\mathrm{PbCl_2}$ precipitate. Thus, the concept map utilizing the stoichiometric ratios is:

so the volume of lead (II) nitrate that reacted is calculated as:

$$0.123\ L\ \mathrm{NaCl\,solution} \times \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}} \times \frac{1\ mol\ \mathrm{Pb(NO_3)_2}}{2\ mol\ \mathrm{NaCl}} \times \frac{1\ L\ \mathrm{Pb(NO_3)_2\ solution}}{1.5\ mol\ \mathrm{Pb(NO_3)_2}} = 0.041\ \mathrm{Pb(NO_3)_2}$$
$$L\ \mathrm{solution}$$

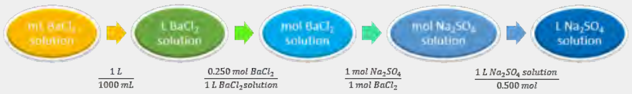
This volume makes intuitive sense for two reasons: (1) the number of moles of $\mathrm{Pb(NO_3)_2}$ required is half of the number of moles of NaCl, based off of the stoichiometry in the balanced reaction (Equation 9.8.1); (2) the concentration of $\mathrm{Pb(NO_3)_2}$ solution is 50% greater than the NaCl solution, so less volume is needed.

Example 9.8.1

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate all $\mathrm{Ba^{2+}}$ in the solution?

Solution

Solutions to Example 13.8.1

Steps for Problem Solving	Example 9.8.1
Identify the "given" information and what the problem is asking you to "find."	Given: 275 mL $\mathrm{BaCl_2}$ 0.250 M $\mathrm{BaCl_2}$ or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M $\mathrm{Na_2SO_4}$ or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume $\mathrm{Na_2SO_4}$ solution.
Set up and balance the chemical equation.	$\mathrm{Na_2SO_4(aq) + BaCl_2(aq) \longrightarrow BaSO_4(s) + \underline{2}NaCl(aq)}$ An insoluble product is formed after the reaction.
List other known quantities.	1 mol of $\mathrm{Na_2SO_4}$ to 1 mol $\mathrm{BaCl_2}$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$275\ \cancel{mL\ BaCl_2\ solution} \times \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \times \frac{0.250\ \cancel{mol\ BaCl_2}}{1\ \cancel{L\ BaCl_2\ solution}} \times \frac{1\ \cancel{mol\ Na_2}}{1\ \cancel{mol\ Ba}}$ = 0.1375 L sodium sulfate

Steps for Problem Solving	Example 9.8.1
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than barium chloride. Also, the units are correct.

? Exercise 9.8.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

Answer

0.250 L LiOH solution

9.9: Freezing Point Depression and Boiling Point Elevation

Learning Objectives

- Explain what the term "colligative" means, and list the colligative properties.
- Indicate what happens to the boiling point and the freezing point of a solvent when a solute is added to it.
- Calculate boiling point elevations and freezing point depressions for a solution.

People who live in colder climates have seen trucks put salt on the roads when snow or ice is forecast. Why is this done? As a result of the information you explore in this section, you will understand why these events occur. You will also learn to calculate exactly how much of an effect a specific solute can have on the boiling point or freezing point of a solution.

The example given in the introduction is an example of a colligative property. **Colligative properties** are properties that differ based on the concentration of solute in a solvent, but not on the type of solute. What this means for the example above is that people in colder climates do not necessarily need salt to get the same effect on the roads—any solute will work. However, the higher the concentration of solute, the more these properties will change.

Boiling Point Elevation

Water boils at 100^oC at $1\; \text{atm}$ of pressure, but a solution of saltwater does not . When table salt is added to water, the resulting solution has a higher boiling point than the water did by itself. The ions form an attraction with the solvent particles that prevents the water molecules from going into the gas phase. Therefore, the saltwater solution will not boil at 100^oC. In order for the saltwater solution to boil, the temperature must be raised about 100^oC. This is true for any solute added to a solvent; the boiling point will be higher than the boiling point of the pure solvent (without the solute). In other words, when anything is dissolved in water, the solution will boil at a higher temperature than pure water would.

The boiling point elevation due to the presence of a solute is also a colligative property. That is, the amount of change in the boiling point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20 \: \text{m}$solution of table salt and a $0.20 \: \text{m}$solution of hydrochloric acid would have the same effect on the boiling point.

Freezing Point Depression

The effect of adding a solute to a solvent has the opposite effect on the freezing point of a solution as it does on the boiling point. A solution will have a lower freezing point than a pure solvent. The **freezing point** is the temperature at which the liquid changes to a solid. At a given temperature, if a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from going into the solid phase. The solute-solvent interactions require the temperature to decrease further in order to solidify the solution. A common example is found when salt is used on icy roadways. Salt is put on roads so that the water on the roads will not freeze at the normal 0^oC but at a lower temperature, as low as -9^oC. The de-icing of planes is another common example of freezing point depression in action. A number of solutions are used, but commonly a solution such as ethylene glycol, or a less toxic monopropylene glycol, is used to de-ice an aircraft. The aircrafts are sprayed with the solution when the temperature is predicted to drop below the freezing point. The **freezing point depression** is the difference in the freezing points of the solution from the pure solvent. This is true for any solute added to a solvent; the freezing point of the solution will be lower than the freezing point of the pure solvent (without the solute). Thus, when anything is dissolved in water, the solution will freeze at a lower temperature than pure water would.

The freezing point depression due to the presence of a solute is also a colligative property. That is, the amount of change in the freezing point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20 \: \text{m}$solution of table salt and a $0.20 \: \text{m}$solution of hydrochloric acid would have the same effect on the freezing point.

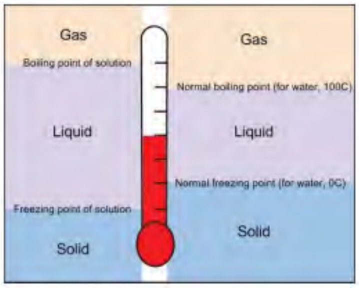


Figure 9.9.1: Comparison of boiling and freezing points of a pure liquid (right side) with a solution (left side).

Comparing the Freezing and Boiling Point of Solutions

Recall that covalent and ionic compounds do not dissolve in the same way. Ionic compounds break up into cations and anions when they dissolve. Covalent compounds typically do not break up. For example a sugar/water solution stays as sugar and water, with the sugar molecules staying as molecules. Remember that colligative properties are due to the number of solute particles in the solution. Adding 10 molecules of sugar to a solvent will produce 10 solute particles in the solution. When the solute is ionic, such as NaCl however, adding 10 formulas of solute to the solution will produce 20 ions (solute particles) in the solution. Therefore, adding enough NaCl solute to a solvent to produce a $0.20\ \mathrm{m}$ solution will have twice the effect of adding enough sugar to a solvent to produce a $0.20\ \mathrm{m}$ solution. Colligative properties depend on the number of solute particles in the solution.

"i" is the number of particles that the solute will dissociate into upon mixing with the solvent. For example, sodium chloride, NaCl, will dissociate into two ions so for NaCl, $i=2$; for lithium nitrate, $\mathrm{LiNO_3}$, $i=2$; and for calcium chloride, $\mathrm{CaCl_2}$, $i=3$. For covalent compounds, i is always equal to 1.

By knowing the *molality* of a solution and the number of particles a compound will dissolve to form, it is possible to predict which solution in a group will have the lowest freezing point. To compare the boiling or freezing points of solutions, follow these general steps:

1. Label each solute as ionic or covalent.
2. If the solute is ionic, determine the number of *ions* in the formula. Be careful to look for polyatomic ions.
3. Multiply the original molality (m) of the solution by the number of particles formed when the solution dissolves. This will give you the total concentration of particles dissolved.
4. Compare these values. The higher total concentration will result in a higher boiling point and a lower freezing point.

✓ Example 9.9.1

Rank the following solutions in water in order of *increasing* (lowest to highest) freezing point:

- $0.1\ \mathrm{m}\ \mathrm{NaCl}$
- $0.1\ \mathrm{m}\ \mathrm{C_6H_{12}O_6}$
- $0.1\ \mathrm{m}\ \mathrm{CaI_2}$

Solution

To compare freezing points, we need to know the total concentration of all particles when the solute has been dissolved.

- $0.1\ \mathrm{m}\ \mathrm{NaCl}$: This compound is ionic (metal with nonmetal), and will dissolve into 2 parts. The total final concentration is: $(0.1\ \mathrm{m})\,(2) = 0.2\ \mathrm{m}$
- $0.1\ \mathrm{m}\ \mathrm{C_6H_{12}O_6}$: This compound is covalent (nonmetal with nonmetal), and will stay as 1 part. The total final concentration is: $(0.1\ \mathrm{m})\,(1) = 0.1\ \mathrm{m}$
- $0.1\ \mathrm{m}\ \mathrm{CaI_2}$: This compound is ionic (metal with nonmetal), and will dissolve into 3 parts. The total final concentration is: $(0.1\ \mathrm{m})\,(3) = 0.3\ \mathrm{m}$

Remember, the greater the concentration of particles, the lower the freezing point will be. $0.1\ \mathrm{m}\ \mathrm{CaI_2}$ will have the lowest freezing point, followed by $0.1\ \mathrm{m}\ \mathrm{NaCl}$, and the highest of the three solutions will be $0.1\ \mathrm{m}\ \mathrm{C_6H_{12}O_6}$, but all three of them will have a lower freezing point than pure water.

The boiling point of a solution is higher than the boiling point of a pure solvent, and the freezing point of a solution is lower than the freezing point of a pure solvent. However, the amount to which the boiling point increases or the freezing point decreases depends on the amount of solute that is added to the solvent. A mathematical equation is used to calculate the boiling point elevation or the freezing point depression.

The boiling point elevation is the amount that the boiling point temperature *increases* compared to the original solvent. For example, the boiling point of pure water at 1.0 atm is 100°C while the boiling point of a 2% saltwater solution is about 102°C. Therefore, the boiling point elevation would be 2°C. The freezing point depression is the amount that the freezing temperature *decreases*.

Both the boiling point elevation and the freezing point depression are related to the molality of the solution. Looking at the formula for the boiling point elevation and freezing point depression, we see similarities between the two. The equation used to calculate the increase in the boiling point is:

$$\Delta T_b = k_b \cdot \text{m} \cdot i \tag{9.9.1}$$

Where:

- $\Delta T_b =$ the amount the boiling point increases.
- $k_b =$ the boiling point elevation constant which depends on the solvent (for water, this number is $0.515^{\circ}\text{C}/\text{m}$).
- $\text{m} =$ the molality of the solution.
- $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1).

The following equation is used to calculate the decrease in the freezing point:

$$\Delta T_f = k_f \cdot \text{m} \cdot i \tag{9.9.2}$$

Where:

- $\Delta T_f =$ the amount the freezing temperature decreases.
- $k_f =$ the freezing point depression constant which depends on the solvent (for water, this number is $1.86^{\circ}\text{C}/\text{m}$).
- $\text{m} =$ the molality of the solution.
- $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1).

✓ Example 9.9.2: Adding Antifreeze to Protein Engines

Antifreeze is used in automobile radiators to keep the coolant from freezing. In geographical areas where winter temperatures go below the freezing point of water, using pure water as the coolant could allow the water to freeze. Since water expands when it freezes, freezing coolant could crack engine blocks, radiators, and coolant lines. The main component in antifreeze is ethylene glycol, $\text{C}_2\text{H}_4(\text{OH})_2$. What is the concentration of ethylene glycol in a solution of water, in molality, if the freezing point dropped by 2.64°C? The freezing point constant, k_f, for water is $1.86^{\circ}\text{C}/\text{m}$.

Solution

Use the equation for freezing point depression of solution (Equation 9.9.2):

$$\Delta T_f = k_f \cdot \text{m} \cdot i$$

Substituting in the appropriate values we get:

$$2.64^{\circ}\text{C} = (1.86^{\circ}\text{C}/\text{m})\,(\text{m})\,(1)$$

Solve for m by dividing both sides by $1.86^{\circ}\text{C}/\text{m}$.

$$\text{m} = 1.42$$

✓ Example 9.9.3: Adding Salt to Elevate Boiling Temperature

A solution of 10.0 g of sodium chloride is added to 100.0 g of water in an attempt to elevate the boiling point. What is the boiling point of the solution? k_b for water is $0.52^{\circ}\text{C}/\text{m}$.

Solution

Use the equation for boiling point elevation of solution (Equation 9.9.1):

$$\Delta T_b = k_b \cdot \text{m} \cdot i$$

We need to be able to substitute each variable into this equation.

- $k_b = 0.52^{\circ}\text{C}/\text{m}$
- m: We must solve for this using stoichiometry. Given: 10.0 g NaCl and $100.0\text{ g H}_2\text{O}$ Find: $\text{mol NaCl}/\text{kg H}_2\text{O}$. Ratios: molar mass of NaCl, $1000\text{ g} = 1\text{ kg}$

$$\frac{10.0\ \text{ g}\ \cancel{\text{NaCl}}}{100.0\ \text{ g}\ \cancel{\text{H}_2\text{O}}} \cdot \frac{1\text{ mol NaCl}}{58.45\ \text{ g}\ \cancel{\text{NaCl}}} \cdot \frac{1000\ \text{ g}\ \cancel{\text{H}_2\text{O}}}{1\text{ kg H}_2\text{O}} = 1.71\text{ m}$$

- For NaCl, $i = 2$

Substitute these values into the equation $\Delta T_b = k_b \cdot \text{m} \cdot i$. We get:

$$\Delta T_b = \left(0.52 \frac{^{\circ}\text{C}}{\cancel{\text{m}}}\right)(1.71\ \cancel{\text{m}})(2) = 1.78^{\circ}\text{C}$$

Water normally boils at 100°C, but our calculation shows that the boiling point increased by 1.78°C. Our new boiling point is 101.78°C.

Note: Since sea water contains roughly 28.0 g of NaCl per liter, this saltwater solution is approximately **four times** more concentrated than sea water (all for a 2° C rise of boiling temperature).

Summary

- Colligative properties are properties that are due only to the number of particles in solution, and are not related to the chemical properties of the solute.
- Boiling points of solutions are higher than the boiling points of the pure solvents.
- Freezing points of solutions are lower than the freezing points of the pure solvents.
- Ionic compounds split into ions when they dissolve, forming more particles. Covalent compounds stay as complete molecules when they dissolve.

Vocabulary

- **Colligative property** - A property that is due only to the number of particles in solution, and not the type of the solute.
- **Boiling point elevation** - The amount that the boiling point of a solution increases from the boiling point of the pure solvent.
- **Freezing point depression** - The amount that the freezing point of a solution decreases from the boiling point of the pure solvent.

9.10: Osmosis

Learning Objectives

- Explain the following laws within the Ideal Gas Law

Before we introduce the final colligative property, we need to present a new concept. A **semipermeable membrane** is a thin membrane that will pass certain small molecules, but not others. A thin sheet of cellophane, for example, acts as a semipermeable membrane. Consider the system in Figure 9.10.1.

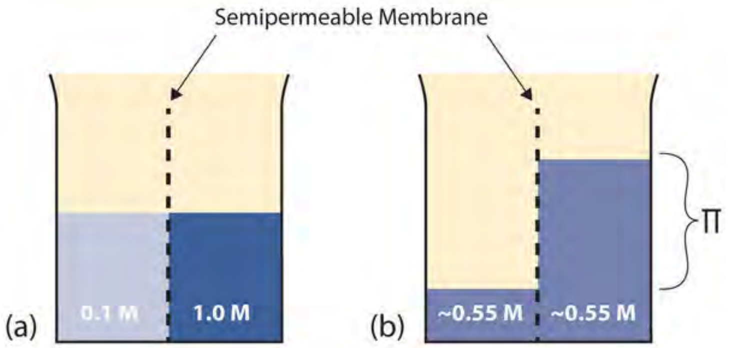


Figure 9.10.1: Osmosis. (a) Two solutions of differing concentrations are placed on either side of a semipermeable membrane. (b) When osmosis occurs, solvent molecules selectively pass through the membrane from the dilute solution to the concentrated solution, diluting it until the two concentrations are the same. The pressure exerted by the different height of the solution on the right is called the osmotic pressure. (CC BY-SA-NC 3.0; anonymous)

a. A semipermeable membrane separates two solutions having the different concentrations marked. Curiously, this situation is not stable; there is a tendency for water molecules to move from the dilute side (on the left) to the concentrated side (on the right) until the concentrations are equalized, as in Figure 9.10.1b
b. This tendency is called **osmosis**. In osmosis, the solute remains in its original side of the system; only solvent molecules move through the semipermeable membrane. In the end, the two sides of the system will have different volumes. Because a column of liquid exerts a pressure, there is a pressure difference (Π) on the two sides of the system that is proportional to the height of the taller column. This pressure difference is called the **osmotic pressure**, which is a colligative property.

The osmotic pressure of a solution is easy to calculate:

\[\Pi = MRT \nonumber \]

where Π is the osmotic pressure of a solution, M is the molarity of the solution, R is the ideal gas law constant, and T is the absolute temperature. This equation is reminiscent of the ideal gas law we considered in Chapter 6.

Example 9.10.5: Osmotic Pressure

What is the osmotic pressure of a 0.333 M solution of $C_6H_{12}O_6$ at 25°C?

Solution

First we need to convert our temperature to kelvins:

$$T = 25 + 273 = 298\ \text{K}$$

Now we can substitute into the equation for osmotic pressure, recalling the value for R:

$$\Pi = (0.333 M)\left(0.08205 \frac{L.atm}{mol.K}\right)(298 K)$$

The units may not make sense until we realize that molarity is defined as moles per liter:

$$\Pi = \left(0.333 \frac{mol}{L}\right)\left(0.08205 \frac{L.atm}{mol.K}\right)(298 K)$$

Now we see that the moles, liters, and kelvins cancel, leaving atmospheres, which is a unit of pressure. Solving,

$$\Pi = 8.14\, atm$$

This is a substantial pressure! It is the equivalent of a column of water 84 m tall.

? Exercise 9.10.5

What is the osmotic pressure of a 0.0522 M solution of $C_{12}H_{22}O_{11}$ at 55°C?

Answer

1.40 atm

Osmotic pressure is important in biological systems because cell walls are semipermeable membranes. In particular, when a person is receiving intravenous (IV) fluids, the osmotic pressure of the fluid needs to be approximately the same as blood serum to avoid any negative consequences. Figure 9.10.3 shows three red blood cells:

- A healthy red blood cell.
- A red blood cell that has been exposed to a lower concentration than normal blood serum (a *hypotonic* solution); the cell has plumped up as solvent moves into the cell to dilute the solutes inside.
- A red blood cell exposed to a higher concentration than normal blood serum (*hypertonic*); water leaves the red blood cell, so it collapses onto itself. Only when the solutions inside and outside the cell are the same (*isotonic*) will the red blood cell be able to do its job.

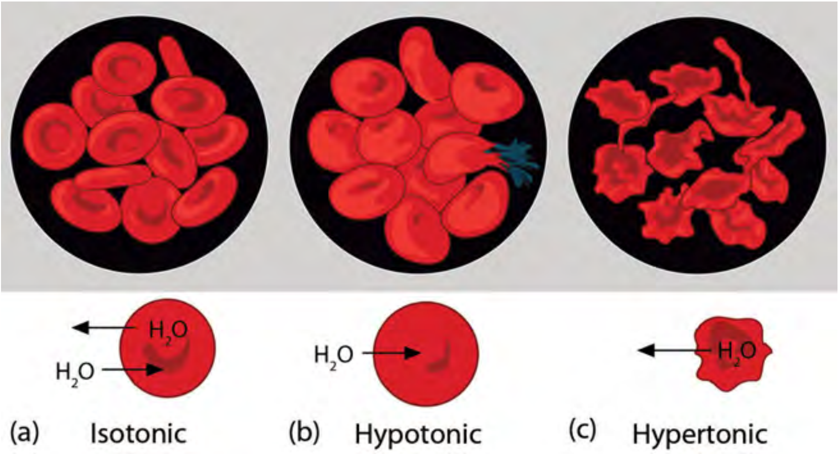


Figure 9.10.3: Osmotic Pressure and Red Blood Cells. (a) This is what a normal red blood cell looks like. (b) When a red blood cell is exposed to a hypotonic solution, solvent goes through the cell membrane and dilutes the inside of the cell. (c) When a red blood cell is exposed to a hypertonic solution, solvent goes from the cell to the surrounding solution, diluting the hypertonic solution and collapsing the cell. Neither of these last two cases is desirable, so IV solutions must be isotonic with blood serum to not cause deleterious effects. (Public Domain; Mariana Ruiz Villareal)

Osmotic pressure is also the reason you should not drink seawater if you're stranded on a lifeboat in the ocean; seawater has a higher osmotic pressure than most of the fluids in your body. You *can* drink the water, but ingesting it will pull water out of your cells as osmosis works to dilute the seawater. Ironically, your cells will die of thirst, and you will also die. (It is okay to drink the water if you are stranded on a body of freshwater, at least from an osmotic pressure perspective.) Osmotic pressure is also thought to be important—in addition to capillary action—in getting water to the tops of tall trees.

Summary

- Osmotic pressure is caused by concentration differences between solutions separated by a semipermeable membrane, and is an important biological consideration.

CHAPTER OVERVIEW

10: Kinetics and Equilibrium

10.1: Chemical Reaction Rate

Drag racing is a sport that involves two cars starting from a dead stop, and driving as fast as they can down a quarter-mile strip. At the end of the strip are timers that determine both elapsed time (how long it took for the cars to cover the quarter mile) and top speed (how fast they were going as they went through the timer chute). Both pieces of data are important. One car may accelerate faster and get ahead that way, while the other car may be slower off the line, but can get up to a higher top speed at the end of the run.

Chemical Reaction Rate

Chemical reactions vary widely in terms of the speed with which they occur. Some reactions occur very quickly. If a lit match is brought in contact with lighter fluid or another flammable liquid, it erupts into flames instantly and burns fast. Other reactions occur very slowly. A container of milk in the refrigerator will be good to drink for weeks before it begins to turn sour. Millions of years were required for dead plants under Earth's surface to accumulate and eventually turn into fossil fuels such as coal and oil.

Chemists need to be concerned with the rates at which chemical reactions occur. Rate is another word for speed. If a sprinter takes 11.0 s to run a 100 m dash, his rate is given by the distance traveled divided by the time.

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100 \text{ m}}{11.0 \text{ s}} = 9.09 \text{ m/s}$$

The sprinter's average running rate for the race is 9.09 m/s We say that it is his average rate because he did not run at that speed for the entire race. At the very beginning of the race, while coming from a standstill, his rate must be slower until he is able to get up to his top speed. His top speed must then be greater than 9.09 m/s so that, taken over the entire race, the average ends up at 9.09 m/s

Figure 10.1.1: Runner. (CC BY-NC; CK-12)

Chemical reactions can't be measured in units of meters per second, as that would not make any sense. A **reaction rate** is the change in concentration of a reactant or product with time. Suppose that a simple reaction were to take place in which a 1.00 M aqueous solution of substance A was converted to substance B.

$$\text{A}\,(aq) \rightarrow \text{B}\,(aq)$$

Suppose that after 20.0 seconds, the concentration of A had dropped from 1.00 M to 0.72 M as A was slowly being converted to B. We can express the rate of this reaction as the change in concentration of A divided by time.

$$\text{rate} = -\frac{\Delta\,[\text{A}]}{\Delta t} = -\frac{[\text{A}]_\text{final} - [\text{A}]_\text{initial}}{\Delta t}$$

A bracket around a symbol or formula means the concentration in molarity of that substance. The change in concentration of A is its final concentration minus its initial concentration. Because the concentration of A is decreasing over time, the negative sign is used. Thus, the rate for the reaction is positive and the units are molarity per second or M/s.

$$\text{rate} = -\frac{0.72 \text{ M} - 1.00 \text{ M}}{20.0 \text{ s}} = -\frac{-0.28 \text{ M}}{20.0 \text{ s}} = 0.014 \text{ M/s}$$

The molarity of A decreases by an average rate of 0.014 M every second. In summary, the rate of a chemical reaction is measured by the change in concentration over time for a reactant or product. The unit of measurement for a reaction rate is molarity per second (M/s).

Summary

- Chemists need to be concerned with the rates at which chemical reactions occur.
- The reaction rate indicates how fast the reaction proceeds.

10.2: Collision Theory

Car damage can be very expensive, especially if the driver who did the damage does not have car insurance. Many people have had the experience of backing up while parallel parking and hearing that "bump". Fortunately, cars are often not going fast enough to do any damage. Every once in a while, though, there is a rearrangement of the body parts of a car when it is hit with sufficient speed; then things need to be fixed.

Collision Theory

The behavior of the atoms, molecules, or ions that comprise the reactants is responsible for the rates of a given chemical reaction. **Collision theory** is a set of principles that states that the reacting particles can form products when they collide with one another, provided those collisions have enough kinetic energy and the correct orientation. Particles that lack the necessary kinetic energy may collide, but the particles will simply bounce off one another unchanged. The figure below illustrates the difference. In the first collision, the particles bounce off one another, and no rearrangement of atoms has occurred. The second collision occurs with greater kinetic energy, and so the bond between the two red atoms breaks. One red atom bonds with the other molecule as one product, while the single red atom is the other product. The first collision is called an **ineffective collision**, while the second collision is called an **effective collision**.

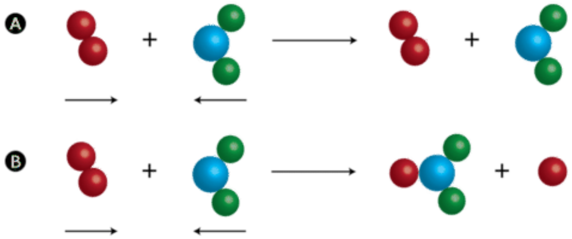


Figure 10.2.1: An ineffective collision (A) is one that does not result in product formation. An effective collision (B) is one in which chemical bonds are broken and a product is formed. (CC BY-NC; CK-12)

Summary

- Collision theory explains how materials can collide and become new materials.
- Effective collisions result in product formation.
- Ineffective collisions do not result in product formation.

10.3: Activation Energy

The sight of fireworks cascading across the night sky is a hallmark of special occasions. These materials, invented hundreds of years ago, can be dangerous if not handled properly. The chemicals do not react until the fuse burns down and heat is applied to the system. Then, the rocket is launched and explodes high in the sky.

Activation Energy

Figure 10.3.1: Calcium metal stored in an argon atmosphere. (CC BY-NC; CK-12)

Why do some chemical reactions occur readily while others require input of heat in order to take place? If we mix metallic sodium with water, a reaction occurs immediately, releasing a great deal of heat (enough to ignite the hydrogen gas that is formed). Group II metals, such as calcium, react at a much slower rate. Unlike the extremely vigorous reaction with sodium, the reaction with calcium is slow enough that we can trap the hydrogen gas released.

Supplying reactant particles with energy causes the bonds between the atoms to vibrate with a greater frequency. This increase in vibrational energy makes a chemical bond more likely to break, and a chemical reaction more likely to occur, when those particles collide with other particles. The **activation energy** for a reaction is the minimum energy that colliding particles must have in order to undergo a reaction. Some reactions occur readily at room temperature because the reacting particles already have the requisite activation energy at that temperature. Other reactions only occur when heated because the particles do not have enough energy unless an external source of heat provides the particles with more kinetic energy.

Summary

10.4: Factors Affecting Reaction Rate

Driving on a crowded freeway can be stressful. Lots of cars, drivers who aren't paying attention, people who speed or who drive too slow—the chances of a collision are rather high. A lot of cars in a particular amount of space means a high car concentration and many opportunities for unwanted connections with other cars.

Factors Affecting Reaction Rates

According to their nature, some reactions occur very quickly, while others are very slow. However, certain changes in the reacting conditions can have an effect on the rate of a given chemical reaction. Collision theory can be utilized to explain these rate effects.

Concentration

An increase in the concentration of one or more reacting substances results in an increase in the rate of reaction. When more particles are present in a given amount of space, a greater number of collisions will naturally occur between those particles. Since the rate of a reaction is dependent on the number of collisions occurring between reactants, the rate increases as the concentration increases.

Pressure

When the pressure of a gas is increased, its particles are forced closer together, decreasing the amount of empty space between the particles. Therefore, an increase in the pressure of a gas is also an increase in the concentration of the gas. For gaseous reactions, an increase in pressure increases the rate of reaction due to a greater number of collisions between reacting particles.

Surface Area

A large log placed in a fire will burn relatively slowly. If the same mass of wood were added to the fire in the form of small twigs, the twigs would burn much more quickly. This is because the twigs provide a greater surface area than the log does. An increase in the surface area of a reactant increases the rate of a reaction. Surface area is larger when a given amount of a solid is present as smaller particles. A powdered reactant has a greater surface area than the same reactant as a solid chunk. In order to increase the surface area of a substance, it may be ground into smaller particles or dissolved into a liquid. In solution, the dissolved particles are separated from each other and will react more quickly with other reactants.

Temperature

Raising the temperature of a chemical reaction usually results in a higher rate of reaction. When the reactant particles are heated, they move faster and faster. This results in a greater frequency of collisions. A more important effect of the temperature increase is that the collisions occur with a greater force, and are thus more likely to surmount the activation energy barrier and go on to form products. Increasing the temperature of a reaction increases the number of effective collisions between reacting particles, so the reaction rate increases.

Summary

- Factors that affect (and generally increase) reaction rate are:
  - Concentration of reactants
  - Pressure (in the case of a gas)
  - Surface area
  - Temperature

10.5: Catalysts

Gasoline-powered vehicles emit various harmful materials. Nitrogen oxides are formed when atmospheric nitrogen reacts with oxygen at the high temperatures found in a car engine. Carbon monoxide is a by-product of the incomplete combustion of hydrocarbons. Evaporated and unused fuel releases volatile hydrocarbons into the atmosphere to help form smog. The presence of a catalytic converter in the exhaust system of the car causes these materials to react and be changed into less harmful products.

Catalysts

Sometimes a particular substance added to a chemical reaction will cause that reaction to undergo a dramatic increase in rate. Hydrogen peroxide is used as a disinfectant for scrapes and cuts, and is found in many medicine cabinets as a 3% aqueous solution. Hydrogen peroxide naturally decomposes to produce water and oxygen gas, but the reaction is very slow. A bottle of hydrogen peroxide will last for several years before it needs to be replaced. However, the addition of just a small amount of manganese (IV) oxide to hydrogen peroxide will cause it to decompose completely in just a matter of minutes. A **catalyst** is a substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction. After the reaction occurs, a catalyst returns to its original state; so catalysts can be used over and over again. Because it is neither a reactant nor a product, a catalyst is shown in a chemical equation by being written above the yield arrow.

$$\mathrm{H_2O_2}\,(aq) \overset{\mathrm{MnO_2}}{\rightarrow} \mathrm{2H_2O}\,(l) + \mathrm{O_2}\,(g)$$

A catalyst works by changing the specific way in which the reaction occurs, called its mechanism. The important outcome from the use of a catalyst is that the overall activation energy of the reaction is lowered (see figure below). With a lower activation energy barrier, a greater percentage of reactant molecules are able to have effective collisions, and the reaction rate increases.

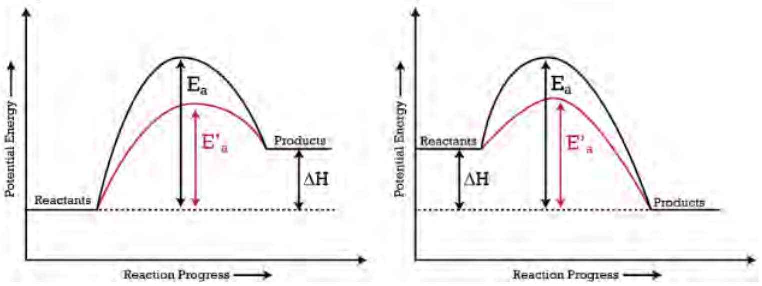


Figure 10.5.1: The addition of a catalyst to a reaction lowers the activation energy, increasing the rate of the reaction. The activation energy of the uncatalyzed reaction is shown by E_a, while the catalyzed reaction is shown by E'_a. The heat of reaction (ΔH) is unchanged by the presence of the catalyst. (CC BY-NC; CK-12)

Catalysts are an extremely important part of many chemical reactions. Enzymes in your body acts as nature's catalysts, allowing important biochemical reactions to occur at reasonable rates. Chemical companies constantly search for new and better catalysts to make reactions go faster and thus make the company more profitable.

Summary

- A catalyst is a substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction.
- Catalysts are an important part of many chemical reactions.

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10.6: Rate Law and Specific Rate Constant

Where are people moving from and where are they moving to? How fast is the population changing in different areas? These are important considerations for those individuals or companies who decide where to build schools or hospitals, or where to open new businesses. If an area is growing rapidly, action needs to be taken quickly to accommodate the growth. Rates of change affect many decisions.

Rate Law and Specific Rate Constant

Consider a simple chemical reaction in which reactant A is converted into product B, according to the equation below.

$$\mathrm{A} \rightarrow \mathrm{B}$$

The rate of reaction is given by the change in concentration of A as a function of time. The rate of disappearance of A is also proportional to the concentration of A.

$$-\frac{\Delta\,[\mathrm{A}]}{\Delta t} \propto [\mathrm{A}]$$

Since the rate of a reaction generally depends upon collision frequency, it stands to reason that as the concentration of A increases, the rate of reaction increases. Likewise, as the concentration of A decreases, the rate of reaction decreases. The expression for the rate of the reaction can be shown as follows:

$$\text{rate} = -\frac{\Delta\,[\mathrm{A}]}{\Delta t} \quad \text{or} \quad \text{rate} = k\,[\mathrm{A}]$$

The proportionality between the rate and $[\mathrm{A}]$ becomes an equal sign by the insertion of a constant (k). A **rate law** is an expression showing the relationship of the reaction rate to the concentrations of each reactant. The **specific rate constant** $(\boldsymbol{k})$ is the proportionality constant relating the rate of the reaction to the concentrations of reactants. The rate law and the specific rate constant for any chemical reaction must be determined experimentally. The value of the rate constant is temperature dependent. A large value of the rate constant means that the reaction is relatively fast, while a small value of the rate constant means that the reaction is relatively slow.

Summary

- A rate law is an expression showing the relationship of the reaction rate to the concentrations of each reactant.
- The specific rate constant (k) is the proportionality constant relating the rate of the reaction to the concentrations of reactants.
- The rate law and the specific rate constant for any chemical reaction must be determined experimentally.

10.7: Order of Reaction

Forest fires cause extensive damage; both plant and animal life are harmed. The severity of a forest fire depends on how much plant life is available to burn—the more available dry plant material, the more serious the fire, and the more rapidly it will spread.

Order of Reaction

In the reaction $\text{A} \rightarrow \text{B}$, the rate of the reaction is directly proportional to the concentration of A raised to the first power. That is to say, $[\text{A}] = [\text{A}]^1$. A **first-order reaction** is a reaction in which the rate is proportional to the concentration of only one reactant. As a first-order reaction proceeds, the rate of reaction decreases because the concentration of the reactant decreases (see figure below). The graph of concentration versus time is curved. The reaction rate $\left(\frac{\Delta[\text{A}]}{\Delta t}\right)$ can be determined graphically by the slope of a tangent to the curve at any point. The rate of the reaction at the time shown with the red triangle is given by:

$$\text{rate} = -\frac{[\text{A}]_{\text{final}} - [\text{A}]_{\text{initial}}}{\Delta t} = -\frac{0.35\text{ M} - 0.63\text{ M}}{3.0\text{ s} - 1.0\text{ s}} = 0.14\text{ M/s}$$

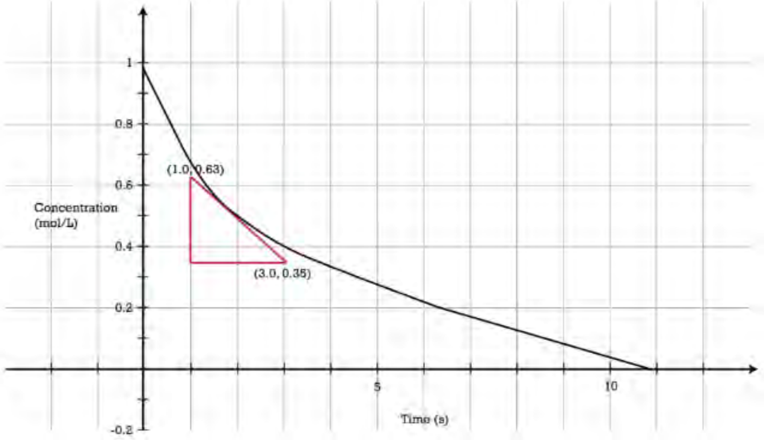

Figure 10.7.1: This graph shows how the concentration of a reactant changes as a reaction proceeds. The rate of the reaction is determined at any point by measuring the slope of a tangent to the curve. (CC BY-NC; CK-12)

The rates of some reactions depend on the concentrations of more than one reactant. Consider a reaction in which a molecule of A collides with a molecule of B to form product C.

$$\text{A} + \text{B} \rightarrow \text{C}$$

Doubling the concentration of A alone would double the reaction rate. Likewise, doubling the concentration of B alone would also double the rate. The rate law must reflect the rate dependence on both reactants.

$$\text{rate} = k[\text{A}][\text{B}]$$

This reaction is said to be first order with respect to A and first order with respect to B. Overall, it is a second-order reaction. The rate law and the order of a reaction must be determined experimentally.

Summary

- A first-order reaction is a reaction in which the rate is proportional to the concentration of only one reactant.
- As a first-order reaction proceeds, the rate of reaction decreases because the concentration of the reactant decreases (see figure below).

10.8: Rate-Determining Step

Airline travel can be very frustrating. Travelers usually have to get to the airport two hours before their flight leaves. They stand in line to check their baggage and get a boarding pass. Then they stand in line for security screening. Finally, travelers wait in line to board the plane. Since there are only so many ticket agents, not everybody can be waited on immediately. The same with the security screen—only so many body scanners are available. And getting on the plane involves going one-by-one down a very narrow aisle to get to a designated seat. All these limits slow down the airline travel process.

Rate-Determining Step

The determination of a reaction mechanism can only be made in the laboratory. When a reaction occurs in a sequence of elementary steps, the overall reaction rate is governed by whichever one of those steps is the slowest. The **rate-determining step** is the slowest step in the sequence of steps in a reaction mechanism. Imagine driving on a one-lane road where it is not possible to pass another vehicle. The rate of flow of traffic on such a road would be dictated by whichever car is traveling at the lowest speed. The rate-determining step is a similar concept to this slow car determining the traffic flow rate—the overall reaction rate is determined by the slowest part of the process.

Decomposition of Hydrogen Peroxide

Recall that a catalyst is a substance that increases the rate of a chemical reaction without being consumed. Catalysts lower the overall activation energy for a reaction by providing an alternative mechanism for the reaction to follow. One such catalyst for the decomposition of hydrogen peroxide is the iodide ion $\left(\mathrm{I}^-\right)$.

$$2\mathrm{H_2O_2}\,(aq) \xrightarrow{\mathrm{I}^-} 2\mathrm{H_2O}\,(l) + \mathrm{O_2}\,(g)$$

By experiment, the rate of reaction is found to be first-order with respect to both $\mathrm{H_2O_2}$ and ceI^-, and second order overall.

$$\text{rate} = k\left[\mathrm{H_2O_2}\right]\left[\mathrm{I}^-\right]$$

The reaction cannot occur in one step corresponding to the overall balanced equation. If it did, the reaction would be second-order with respect to $\mathrm{H_2O_2}$, since the coefficient of the $\mathrm{H_2O_2}$ in the balanced equation is a 2. A reaction mechanism that accounts for the rate law, and for the detection of the IO^- ion as an intermediate, can be constructed. It consists of two bimolecular elementary steps:

Step 1: $\mathrm{H_2O_2}\,(aq) + \mathrm{I}^-\,(aq) \rightarrow \mathrm{H_2O}\,(l) + \mathrm{IO}^-\,(aq)$

Step 2: $\mathrm{H_2O_2}\,(aq) + \mathrm{IO}^-\,(aq) \rightarrow \mathrm{H_2O_2}\,(l) + \mathrm{O_2}\,(g) + \mathrm{I}^-\,(aq)$

If step 2 is the rate-determining step, then the rate law for that step will be the rate law for the overall reaction.

$$\text{rate} = k\left[\mathrm{H_2O_2}\right]\left[\mathrm{I}^-\right]$$

The rate law for the slow step of the proposed mechanism agrees with the overall experimentally determined rate law. The IO^- is present as an intermediate in the reaction. The iodide catalyst also appears in the mechanism. It is consumed in the first elementary step and then is regenerated in the second step. This is the requirement for a catalyst—to not be used up in the reaction.

Summary

- The rate-determining step in a reaction is defined.
- The process for determining the rate-determining step is described.

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10.9: Chemical Equilibrium

A tug of war involves two teams at the ends of a rope. The goal is to pull the other team over a line in the middle. At first, there is a great deal of tension on the rope, but no apparent movement. A bystander might think that there is nothing happening. In reality, there is a great deal of tension on the rope as the two teams pull in opposite directions at the same time.

Chemical Equilibrium

Hydrogen and iodine gases react to form hydrogen iodide, according to the following reaction:

$$\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g)$$

$$\text{Forward reaction:} \quad \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightarrow 2\mathrm{HI}\,(g)$$

$$\text{Reverse reaction:} \quad 2\mathrm{HI}\,(g) \rightarrow \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g)$$

Initially, only the forward reaction occurs because no HI is present. As soon as some HI has formed, it begins to decompose back into $\mathrm{H_2}$ and $\mathrm{I_2}$. Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually, the rate of combination of $\mathrm{H_2}$ and $\mathrm{I_2}$ to produce HI becomes equal to the rate of decomposition of HI into $\mathrm{H_2}$ and $\mathrm{I_2}$. When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance. **Chemical equilibrium** is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.

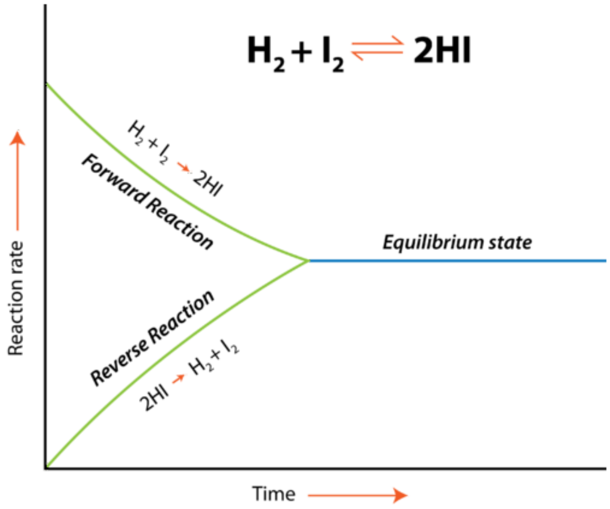


Figure 10.9.1: Equilibrium in reaction: $\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g)$.

Chemical equilibrium can be attained whether the reaction begins with all reactants and no products, all products and no reactants, or some of both. The figure below shows changes in concentration of $\mathrm{H_2}$, $\mathrm{I_2}$, and HI for two different reactions. In the reaction depicted by the graph on the left (A), the reaction begins with only $\mathrm{H_2}$ and $\mathrm{I_2}$ present. There is no HI initially. As the reaction proceeds toward equilibrium, the concentrations of the $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually decrease, while the concentration of the HI gradually increases. When the curve levels out and the concentrations all become constant, equilibrium has been reached. At equilibrium, concentrations of all substances are constant. In reaction B, the process begins with only HI and no $\mathrm{H_2}$ or $\mathrm{I_2}$. In this case, the concentration of HI gradually decreases while the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually increase until equilibrium is again reached. Notice that in both cases, the relative position of equilibrium is the same, as shown by the relative concentrations of reactants and products. The concentration of HI at equilibrium is significantly higher than the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$. This is true whether the reaction began with all reactants or all products. The position of equilibrium is a property of the particular reversible reaction and does not depend upon how equilibrium was achieved.

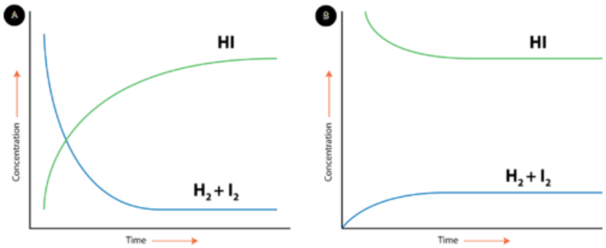


Figure 10.9.2: Equilibrium between reactants and products.

Conditions for Equilibrium

It may be tempting to think that once equilibrium has been reached, the reaction stops. Chemical equilibrium is a dynamic process. The forward and reverse reactions continue to occur even after equilibrium has been reached. However, because the rates of the reactions are the same, there is no change in the relative concentrations of reactants and products for a reaction that is at equilibrium. The conditions and properties of a system at equilibrium are summarized below.

1. The system must be closed, meaning no substances can enter or leave the system.
2. Equilibrium is a dynamic process. Even though we don't necessarily see the reactions, both the forward and reverse reactions are taking place.
3. The rates of the forward and reverse reactions must be equal.
4. The amount of reactants and products do not have to be equal. However, after equilibrium is attained, the amounts of reactants and products will be constant.

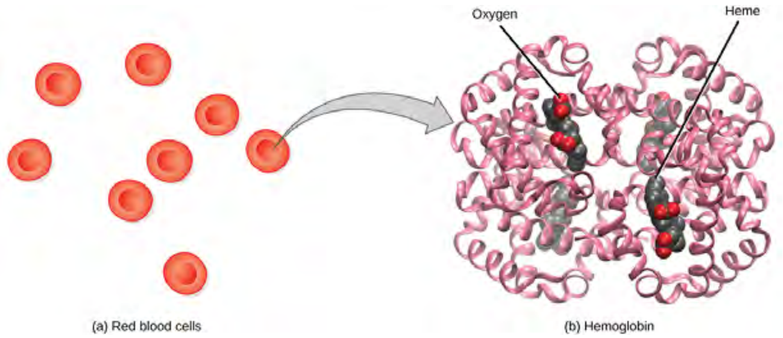
The description of equilibrium in this concept refers primarily to equilibrium between reactants and products in a chemical reaction. Other types of equilibrium include phase equilibrium and solution equilibrium. A **phase equilibrium** occurs when a substance is in equilibrium between two states. For example, a stoppered flask of water attains equilibrium when the rate of evaporation is equal to the rate of condensation. A **solution equilibrium** occurs when a solid substance is in a saturated solution. At this point, the rate of dissolution is equal to the rate of recrystallization. Although these are all different types of transformations, most of the rules regarding equilibrium apply to any situation in which a process occurs reversibly.

Summary

- Chemical equilibrium is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.
- Conditions for chemical equilibrium are listed.
- A phase equilibrium occurs when a substance is in equilibrium between two states.
- A solution equilibrium occurs when a solid substance is in a saturated solution.

10.10: Equilibrium Constant

Red blood cells transport oxygen to tissues so that they can function. In the absence of oxygen, cells cannot carry out their biochemical responsibilities. Oxygen moves to the cells attached to hemoglobin, a protein found in red blood cells.



The protein inside (a) red blood cells that carries oxygen to cells and carbon dioxide to the lungs is (b) hemoglobin. Hemoglobin is made up of four symmetrical subunits and four heme groups. Iron associated with the heme binds oxygen. It is the iron in hemoglobin that gives blood its red color. (CC_BY Openstax)

In cases of carbon monoxide poisoning, CO binds much more strongly to the hemoglobin, blocking oxygen attachment and lowering the amount of oxygen reaching the cells. Treatment involves the patient breathing pure oxygen to displace the carbon monoxide. The equilibrium reaction shown below illustrates the shift toward the right when excess oxygen is added to the system:

$$\mathrm{Hb(CO)_4}\,(aq) + 4\mathrm{O_2}\,(g) \rightleftharpoons \mathrm{Hb(O_2)_4}\,(aq) + 4\mathrm{CO}\,(g)$$

Equilibrium Constant

Consider the hypothetical reversible reaction in which reactants A and B react to form products C and D. This equilibrium can be shown as demonstrated below, where the lower case letters represent the coefficients of each substance.

$$a\mathrm{A} + b\mathrm{B} \rightleftharpoons c\mathrm{C} + d\mathrm{D}$$

As we have established, the rates of the forward and reverse reactions are the same at equilibrium, and so the concentrations of all of the substances are constant. Since that is the case, it stands to reason that a ratio of the concentrations for any given reaction at equilibrium maintains a constant value. The **equilibrium constant** $(\boldsymbol{K}_{\mathrm{eq}})$ is the ratio of the mathematical product of the concentrations of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction. Each concentration is raised to the power of its coefficient in the balanced chemical equation. For the general reaction above, the equilibrium constant expression is written as follows:

$$K_{\mathrm{eq}} = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}$$

The concentrations of each substance, indicated by the square brackets around the formula, are measured in molarity units $(\mathrm{mol/L})$.

The value of the equilibrium constant for any reaction is only determined by experiment. As detailed in the previous section, the position of equilibrium for a given reaction does not depend on the starting concentrations, and so the value of the equilibrium constant is truly constant. It does, however, depend on the temperature of the reaction. This is because equilibrium is defined as a condition resulting from the rates of forward and reverse reactions being equal. If the temperature changes, the corresponding change in those reaction rates will alter the equilibrium constant. For any reaction in which a K_{eq} is given, the temperature should be specified.

Summary

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10.11: Calculations with Equilibrium Constants

Iron is an important component of red blood cells. Patients who have low iron will usually be anemic and have a lower than normal number of red blood cells. One way to assess serum iron concentration is with the use of Ferrozine, a complex organic molecule. Ferrozine forms a product with Fe^{3+}, producing a pink color. In order to determine factors affecting the reaction, we need to measure the equilibrium constant. If the equilibrium does not lie far in the direction of products, precautions need to be taken when using this material to measure iron in serum.

Calculations with Equilibrium Constants

The general value of the equilibrium constant gives us information about whether the reactants or the products are favored at equilibrium. Since the product concentrations are in the numerator of the equilibrium expression, a $K_{\mathrm{eq}} > 1$ means that the products are favored over the reactants. A $K_{\mathrm{eq}} < 1$ means that the reactants are favored over the products.

Though it would often seem that the K_{eq} value would have various units depending on the values of the exponents in the expression, the general rule is that any units are dropped. All K_{eq} values will be reported as having no units.

Example 10.11.1

Equilibrium occurs when nitrogen monoxide gas reacts with oxygen gas to form nitrogen dioxide gas:

$$2\mathrm{NO}\,(g) + \mathrm{O}_2\,(g) \rightleftharpoons 2\mathrm{NO}_2\,(g)$$

At equilibrium at $230^{\mathrm{o}}\mathrm{C}$, the concentrations are measured to be $[\mathrm{NO}] = 0.0542\ \mathrm{M}$, $[\mathrm{O}_2] = 0.127\ \mathrm{M}$, and $[\mathrm{NO}_2] = 15.5\ \mathrm{M}$. Calculate the equilibrium constant at this temperature.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $[\mathrm{NO}] = 0.0542\ \mathrm{M}$
- $[\mathrm{O}_2] = 0.127\ \mathrm{M}$
- $[\mathrm{NO}_2] = 15.5\ \mathrm{M}$

Unknown

The equilibrium expression is first written according to the general form in the text. The equilibrium values are substituted into the expression and the value calculated.

Step 2: Solve.

$$K_{\mathrm{eq}} = \frac{[\mathrm{NO}_2]^2}{[\mathrm{NO}]^2\,[\mathrm{O}_2]}$$

Substituting in the concentrations at equilibrium:

$$K_{\mathrm{eq}} = \frac{(15.5)^2}{(0.0542)^2\,(0.127)} = 6.44 \times 10^5$$

Step 3: Think about your result.

> The equilibrium concentration of the product NO_2 is significantly higher than the concentrations of the reactants NO and O_2. As a result, the K_{eq} value is much larger than 1, an indication that the product is favored at equilibrium.

The equilibrium expression only shows those substances whose concentrations are variable during the reaction. A pure solid or a pure liquid does not have a concentration that will vary during a reaction. Therefore, an equilibrium expression omits pure solids and liquids, and only shows the concentrations of gases and aqueous solutions. The decomposition of mercury (II) oxide can be shown by the following equation, followed by its equilibrium expression.

$$2\text{HgO}\,(s) \rightleftharpoons 2\text{Hg}\,(l) + \text{O}_2\,(g) \qquad K_{eq} = [\text{O}_2]$$

The stoichiometry of an equation can also be used in a calculation of an equilibrium constant. At 40^oC, solid ammonium carbamate decomposes to ammonia and carbon dioxide gases.

$$\text{NH}_4\text{CO}_2\text{NH}_2\,(s) \rightleftharpoons 2\text{NH}_3\,(g) + \text{CO}_2\,(g)$$

At equilibrium, the $[\text{CO}_2]$ is found to be 4.71×10^{-3} M. Can the K_{eq} value be calculated from that information alone? Because the ammonium carbamate is a solid, it is not present in the equilibrium expression.

$$K_{eq} = [\text{NH}_3]^2\,[\text{CO}_2]$$

The stoichiometry of the chemical equation indicates that as the ammonium carbamate decomposes, 2 mol of ammonia gas is produced for every 1 mol of carbon dioxide. Therefore, at equilibrium, the concentration of the ammonia will be twice the concentration of carbon dioxide. So,

$$[\text{NH}_3] = 2 \times \left(4.71 \times 10^{-3}\right) = 9.42 \times 10^{-3}\ \text{M}$$

Substituting these values into the K_{eq} expression:

$$K_{eq} = \left(9.42 \times 10^{-3}\right)^2 \left(4.71 \times 10^{-3}\right) = 4.18 \times 10^{-7}$$

Using Equilibrium Constants

The equilibrium constants are known for a great many reactions. Hydrogen and bromine gases combine to form hydrogen bromide gas. At 730^oC, the equation and K_{eq} are given below.

$$\text{H}_2\,(g) + \text{Br}_2\,(g) \rightleftharpoons 2\text{HBr}\,(g) \qquad K_{eq} = 2.18 \times 10^6$$

A certain reaction is begun with only HBr. When the reaction mixture reaches equilibrium at 730^oC, the concentration of bromine gas is measured to be 0.00243 M What is the concentration of the H_2 and the HBr at equilibrium?

Since the reaction begins with only HBr and the mole ratio of H_2 to Br_2 is 1:1, the concentration of H_2 at equilibrium is also 0.00243 M The equilibrium expression can be rearranged to solve for the concentration of HBr at equilibrium:

$$\begin{aligned} K_{eq} &= \frac{[\text{HBr}]^2}{[\text{H}_2]\,[\text{Br}_2]} \\ [\text{HBr}] &= \sqrt{K_{eq}\,[\text{H}_2]\,[\text{Br}_2]} \\ &= \sqrt{2.18 \times 10^6\,(0.00243)\,(0.00243)} = 3.59\ \text{M} \end{aligned}$$

Since the value of the equilibrium constant is very high, the concentration of HBr is much greater than that of H_2 and Br_2 at equilibrium.

Summary

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10.12: Le Châtelier's Principle

There are some who enjoy going up in an airplane, strapping on a parachute, and diving out the door to free-fall before opening the chute, and dropping to the ground. This stressful activity (so they say) relieves the stress of everyday life. The release of adrenaline caused by this stressful behavior is said to promote a mood enhancement that helps you deal better with other stresses in your daily life.

Le Chatelier's Principle

Chemical equilibrium was studied by French chemist Henri Le Chatelier (1850-1936) and his description of how a system responds to a stress on equilibrium has become known as **Le Chatelier's principle:** When a chemical system that is at equilibrium is disturbed by a stress, the system will respond in order to relieve the stress. Stresses to a chemical system involve changes in the concentration of reactants or products, changes in the temperature of the system, or changes in the pressure of the system. We will discuss each of these stresses separately. The change to the equilibrium position in every case is either a favoring of the forward reaction or a favoring of the reverse reaction. When the forward reaction is favored, the concentrations of products increase, while the concentrations of reactants decrease. When the reverse reaction is favored, the concentrations of products decrease, while the concentrations of reactants increase.

Le Chatelier's Principle

Table 10.12.1

Original Equilibrium	**Favored Reaction**	**Result**
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Forward: $\mathrm{A} \rightarrow \mathrm{B}$	$[\mathrm{A}]$ decreases; $[\mathrm{B}]$ increases
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Reverse: $\mathrm{A} \leftarrow \mathrm{B}$	$[\mathrm{A}]$ increases; $[\mathrm{B}]$ decreases

Figure 10.12.1: Henri Le Chatelier.

Summary

10.13: Effect of Concentration

Phenolphthalein is a chemical that has one structure in a high acid environment, and another structure in a low acid environment. If the hydrogen ion concentration is high, the compound is colorless, but turns red if the hydrogen ion concentration is low. By adding hydrogen ions to the solution, or removing them through a chemical reaction, we can vary the color of the dye.

Effect of Concentration

A change in concentration of one of the substances in an equilibrium system typically involves either the addition or the removal of one of the reactants or products. Consider the Haber-Bosch process for the industrial production of ammonia from nitrogen and hydrogen gases:

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

If the concentration of one substance in a system is increased, the system will respond by favoring the reaction that removes that substance. When more N_2 is added, the forward reaction will be favored because the forward reaction uses up N_2 and converts it to NH_3. The forward reaction speeds up temporarily as a result of the addition of a reactant. The position of equilibrium shifts as more NH_3 is produced. The concentration of NH_3 increases, while the concentrations of N_2 and H_2 decrease. After some time passes, equilibrium is reestablished with new concentrations of all three substances. As can be seen in the figure below, if more N_2 is added, a new equilibrium is achieved by the system. The new concentration of NH_3 is higher because of the favoring of the forward reaction. The new concentration of the H_2 is lower. The concentration of N_2 is higher than in the original equilibrium, but went down slightly following the addition of the N_2 that disturbed the original equilibrium. By responding in this way, the value of the equilibrium constant for the reaction, K_{eq}, does not change as a result of the stress to the system.

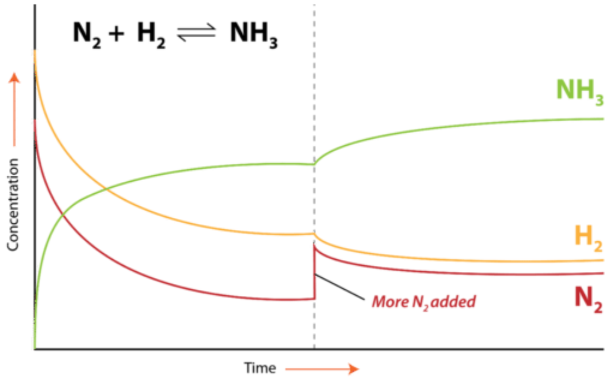

Figure 10.13.1: The Haber-Bosch process is an equilibrium between reactant N_2 and H_2 and product NH_3.

If more NH_3 were added, the reverse reaction would be favored. This "favoring" of a reaction means temporarily speeding up the reaction in that direction until equilibrium is reestablished. Recall that once equilibrium is reestablished, the rates of the forward and reverse reactions are again equal. The addition of NH_3 would result in increased formation of the reactants, N_2 and H_2.

An equilibrium can also be disrupted by the removal of one of the substances. If the concentration of a substance is decreased, the system will respond by favoring the reaction that replaces that substance. In the industrial Haber-Bosch process, NH_3 is removed from the equilibrium system as the reaction proceeds. As a result, the forward reaction is favored so that more NH_3 will be produced. The concentrations of N_2 and H_2 decrease. Continued removal of NH_3 will eventually force the reaction to go to completion until all of the reactants are used up. If either N_2 or H_2 were removed from the equilibrium system, the reverse reaction would be favored, and the concentration of NH_3 would decrease.

The effects of changes in concentration on an equilibrium system, according to Le Chatelier's principle, are summarized in the table below.

Effects of changes in concentration on an equilibrium system, according to Le Chatelier's principle

Table 10.13.1

Stress	**Response**
Addition of reactant	Forward reaction favored.
Addition of product	Reverse reaction favored.

Table 10.13.1

Removal of reactant	Reverse reaction favored.
Removal of product	Forward reaction favored.

Summary

10.14: Effect of Temperature

Carbon monoxide is often thought of as nothing more than a hazardous gas produced from incomplete combustion of carbon products. However, there is a large market for industrially-manufactured carbon monoxide that is used to synthesize most of the acetic acid produced in the world. One reaction that leads to CO formation involves its formation by passing air over excess carbon at high temperatures. The initial product (carbon dioxide) equilibrates with the remaining hot carbon, forming carbon monoxide. At lower temperatures, $\mathrm{CO_2}$ formation is favored, while CO is the predominant product above $800^{\mathrm{o}}\mathrm{C}$.

Effect of Temperature

Increasing or decreasing the temperature of a system at equilibrium is also a stress to the system. The equation for the Haber-Bosch process is written again below, as a thermochemical equation.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g) + 91\text{ kJ}$$

The forward reaction is the exothermic direction: the formation of $\mathrm{NH_3}$ releases heat. The reverse reaction is the endothermic direction: as $\mathrm{NH_3}$ decomposes to $\mathrm{N_2}$ and $\mathrm{H_2}$, heat is absorbed. An increase in the temperature of a system favors the direction of the reaction that absorbs heat, the endothermic direction. Absorption of heat in this case is a relief of the stress provided by the temperature increase. For the Haber-Bosch process, an increase in temperature favors the reverse reaction. The concentration of $\mathrm{NH_3}$ in the system decreases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ increase.

A decrease in the temperature of a system favors the direction of the reaction that releases heat, the exothermic direction. For the Haber-Bosch process, a decrease in temperature favors the forward reaction. The concentration of $\mathrm{NH_3}$ in the system increases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ decrease.

For changes in concentration, the system responds in such a way that the value of the equilibrium constant, K_{eq}, is unchanged. However, a change in temperature shifts the equilibrium, and the K_{eq} value either increases or decreases. As discussed in the previous section, values of K_{eq} are dependent on the temperature. When the temperature of the system for the Haber-Bosch process is increased, the resultant shift in equilibrium toward the reactants means that the K_{eq} value decreases. When the temperature is decreased, the shift in equilibrium toward the products means that the K_{eq} value increases.

Le Chatelier's principle, as related to temperature changes, can be illustrated by the reaction in which dinitrogen tetroxide is in equilibrium with nitrogen dioxide:

$$\mathrm{N_2O_4}\,(g) + \text{heat} \rightleftharpoons 2\mathrm{NO_2}\,(g)$$

Dinitrogen tetroxide $\left(\mathrm{N_2O_4}\right)$ is colorless, while nitrogen dioxide $\left(\mathrm{NO_2}\right)$ is dark brown in color. When $\mathrm{N_2O_4}$ breaks down into $\mathrm{NO_2}$, heat is absorbed according to the forward reaction above. Therefore, an increase in temperature of the system will favor the forward reaction. Conversely, a decrease in temperature will favor the reverse reaction.

Summary

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10.15: Effect of Pressure

The ammonia storage tank in the picture below does two things. First, it stores ammonia at high pressure to minimize the reverse reaction that would lead to less ammonia, and more nitrogen and hydrogen. Secondly, it sends an important message. Ammonia is used to make methamphetamine, a dangerous drug of abuse. Locks and other safety mechanisms built into the tanks help stop the theft of ammonia to be used for this illicit activity.

Figure 10.15.1: An ammonia storage tank.

Effect of Pressure

Changing the pressure of an equilibrium system in which gases are involved is also a stress to the system. A change in the pressure on a liquid or a solid has a negligible effect. We will return again to the equilibrium for the Haber-Bosch process. Imagine the gases are contained in a closed system in which the volume of the system is controlled by an adjustable piston, as shown in the figure below.

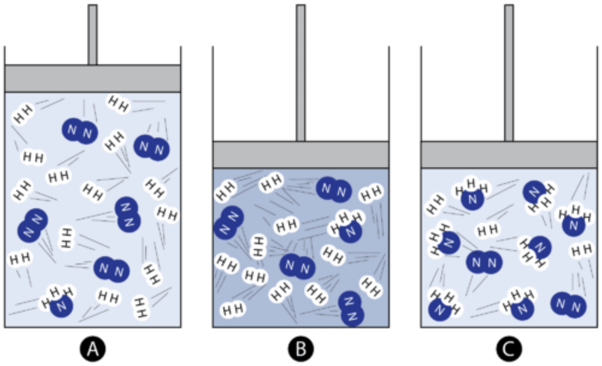

Figure 10.15.2: Effect of pressure on ammonia formation.

On the far left, the reaction system contains primarily N_2 and H_2, with only one molecule of NH_3 present. As the piston is pushed inwards, the pressure of the system increases according to Boyle's Law. This is a stress to the equilibrium. In the middle image, the same number of molecules is now confined in a smaller space and so the pressure has increased. According to Le Chatelier's principle, the system responds in order to relieve the stress. In the image on the right, the forward reaction has been favored, in which one molecule of N_2 combines with three molecules of H_2 to form two molecules of NH_3. The overall result is a decrease in the number of gas molecules in the entire system. This in turn decreases the pressure and provides a relief to the original stress of a pressure increase. An increase in pressure on an equilibrium system favors the reaction which produces fewer total moles of gas. In this case, it is the forward reaction that is favored.

A decrease in pressure on the above system could be achieved by pulling the piston outward, increasing the container volume. The equilibrium would respond by favoring the reverse reaction in which NH_3 decomposes to N_2 and H_2. This is because the overall number of gas molecules would increase, and so would the pressure. A decrease in pressure on an equilibrium system favors the reaction that produces more total moles of gas. This is summarized in the table below.

Table 10.15.1

Stress	Response

Stress	Response
Pressure increase	Reaction produces fewer gas molecules.
Pressure decrease	Reaction produces more gas molecules.

Like changes in concentration, the K_{eq} value for a given reaction is unchanged by a change in pressure.

It is important to remember when analyzing the effect of a pressure change on equilibrium that only gases are affected. If a certain reaction involves liquids or solids, they should be ignored. Calcium carbonate decomposes according to the equilibrium reaction:

$$\mathrm{CaCO_3}\,(s) \rightleftharpoons \mathrm{CaO}\,(s) + \mathrm{O_2}\,(g)$$

Oxygen is the only gas in the system. An increase in the pressure of the system slows the rate of decomposition of $\mathrm{CaCO_3}$ because the reverse reaction is favored. When a system contains equal moles of gas on both sides of the equation, pressure has no effect on the equilibrium position, as in the formation of HCl from $\mathrm{H_2}$ and $\mathrm{Cl_2}$.

$$\mathrm{H_2}\,(g) + \mathrm{Cl_2}\,(g) \rightleftharpoons 2\mathrm{HCl}\,(g)$$

Summary

10.16: Nonreversible Reactions

Fires are a part of life. Some fires clear the land and allow new growth. Other fires provide warmth on a cold night. Unfortunately, many fires are destructive, leaving damage in their wake. All fires leave the environment changed, never to revert back to its original state. The carbon dioxide and water generated by a fire go off into the atmosphere and do not return. The change is permanent and irreversible.

Going to Completion

When one of the products of a reaction is removed from the chemical equilibrium system as soon as it is produced, the reverse reaction cannot establish itself and equilibrium is never reached. Reactions such as these are said to go to completion. These processes are often referred to as **non-reversible reactions**. Reactions which go to completion tend to produce one of three types of products: (1) an insoluble precipitate, (2) a gas, (3) a molecular compound such as water. Examples of these reactions are shown below.

1. Formation of a precipitate:

$$\mathrm{AgNO_3}\left(aq\right)+\mathrm{NaCl}\left(aq\right)\rightarrow\mathrm{NaNO_3}\left(aq\right)+\mathrm{AgCl}\left(s\right)$$

2. Formation of a gas:

$$\mathrm{Mg}\left(s\right)+2\mathrm{HCl}\left(aq\right)\rightarrow\mathrm{MgCl_2}\left(aq\right)+\mathrm{H_2}\left(g\right)$$

3. Formation of water:

$$\mathrm{HCl}\left(aq\right)+\mathrm{NaOH}\left(aq\right)\rightarrow\mathrm{NaCl}\left(aq\right)+\mathrm{H_2O}\left(l\right)$$

If we look at these reactions in more detail, we can see some things that are not apparent the way the equations are written. Looking at the first equation, we do not see a double arrow between reactants and products, because the reaction is considered essentially irreversible. However, if we consider the net ionic equation

$$\mathrm{Ag^+}+\mathrm{Cl^-}\rightarrow\mathrm{AgCl}$$

then the reverse reaction would be:

$$\mathrm{AgCl}\rightarrow\mathrm{Ag^+}+\mathrm{Cl^-}$$

The K_{eq} for the reverse reaction is 1.8×10^{-10}. For all practical purposes, the reaction goes to completion.

Formation of a gas in an open system is essentially irreversible since the gas escapes into the atmosphere. Looking at the activity series, we see that Mg is much higher in the series than hydrogen. So the reaction would be expected to go strongly in the indicated direction.

The third reaction gets a little more complicated. In solution, the reactants HCl and NaOH will be ionized completely, as does the NaCl product. Water exists in an equilibrium with $\mathrm{H^+}$ and $\mathrm{OH^-}$, with the dissociation constant for water being 1×10^{-14}. So, in the solution resulting from the reaction given here, the $\left[\mathrm{H^+}\right]$ is $1\times10^{-7}\ \mathrm{M}$, a very insignificant amount. For all practical purposes, this reaction can be said to go to completion.

Summary

10.17: Le Châtelier's Principle and the Equilibrium Constant

With online banking, management of personal finances can become less complicated in some ways. You can automatically deposit paychecks, pay bills, and designate how much goes into savings or other special accounts each month. If you want to maintain 10% of your bank account in savings, you can set up a program that moves money in and out of the account when you get a paycheck or pay bills. The amount of money in savings will change as the money comes in and out of the bank, but the ratio of savings to checking will always be constant.

Le Chatelier's Principle and the Equilibrium Constant

Occasionally, when students apply Le Chatelier's principle to an equilibrium problem involving a change in concentration, they assume that K_{eq} must change. This seems logical since we talk about "shifting" the equilibrium in one direction or the other. However, K_{eq} is a constant, for a given equilibrium at a given temperature, so it must not change. Here is an example of how this works. Consider the simplified equilibrium below:

$$\mathrm{A} \rightleftharpoons \mathrm{B}$$

Let's say we have a 1.0 liter container. At equilibrium, the following amounts are measured:

$$\begin{aligned} \mathrm{A} &= 0.50 \text{ mol} \\ \mathrm{B} &= 1.0 \text{ mol} \end{aligned}$$

The value of K_{eq} is given by:

$$K_{\mathrm{eq}} = \frac{[\mathrm{B}]}{[\mathrm{A}]} = \frac{1.0 \text{ M}}{0.50 \text{ M}} = 2.0$$

Now we will disturb the equilibrium by adding 0.50 moles of A to the mixture. The equilibrium will shift towards the right, forming more B. Immediately after the addition of A and before any response, we how have 1.0 mol of A and 1.0 mol of B. The equilibrium then shifts in the forward direction. We will introduce a variable (x), which will represent the change in concentrations as the reaction proceeds. Since the mole ratio of A:B is 1:1, as $[\mathrm{A}]$ decreases by the amount x, the $[\mathrm{B}]$ increases by the amount x. We set up an analysis called *ICE*, which stands for Initial, Change, and Equilibrium. The values in the table below represent molar concentrations.

	A	B
Initial	1.0	1.0
Change	$-x$	$+x$
Equilibrium	$1.0-x$	$1.0+x$

At the new equilibrium position, the values for A and B as a function of x can be set equal to the value of the K_{eq}. Then, one can solve for x.

$$K_{\mathrm{eq}} = 2.0 = \frac{[\mathrm{B}]}{[\mathrm{A}]} = \frac{1.0+x}{1.0-x}$$

Solving for x:

$$\begin{aligned} 2.0\,(1.0-x) &= 1.0+x \\ 2.0-2.0x &= 1.0+x \\ 3.0x &= 1.0 \\ x &= 0.33 \end{aligned}$$

This value for x is now plugged back in to the Equilibrium line of the table, and used to find the final concentrations of A and B after the reaction.

$$[\mathrm{A}] = 1.0 - x = 0.67 \text{ M}$$

$$[\mathrm{B}] = 1.0 + x = 1.33 \text{ M}$$

The value of K_{eq} has been maintained since $\frac{1.33}{0.67} = 2.0$. This shows that even though a change in concentration of one of the substances in equilibrium causes a shift in the equilibrium position, the value of the equilibrium constant does not change.

Summary

10.18: Solubility Product Constant (K sp) (Ksp)

At one time, a major analytical technique was gravimetric analysis. Gravimetric analysis involves an ion being precipitated out of solution, purified, and weighed to determine the amount of that ion in the original material. As an example, measurement of Ca^{2+} involved dissolving the sample in water, precipitating the calcium as calcium oxalate, purifying the precipitate, drying it, and weighing the final product. Although this approach can be very accurate (atomic weights for many elements were determined this way), the process is slow, tedious, and prone to a number of errors in technique. Newer methods are now available that measure minute amounts of calcium ions in solution without the long, involved gravimetric approach.

Solubility Product Constant

Ionic compounds have widely differing solubilities. Sodium chloride has a solubility of about $360 \mathrm{~g}$ per liter of water at $25^{\circ}\mathrm{C}$. Salts of alkali metals tend to be quite soluble. On the other end of the spectrum, the solubility of zinc hydroxide is only $4.2 \times 10^{-4} \mathrm{~g/L}$ of water at the same temperature. Many ionic compounds containing hydroxide are relatively insoluble.

Most ionic compounds that are considered to be insoluble will still dissolve to a small extent in water. These "mostly insoluble" compounds are considered to be strong electrolytes, because the portion of the compound that dissolves also dissociates. As an example, silver chloride dissociates to a small extent into silver ions and chloride ions upon being added to water.

$$\mathrm{AgCl}\,(s) \rightleftharpoons \mathrm{Ag}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

The process is written as an equilibrium because the dissociation occurs only to a small extent. Therefore, an equilibrium expression can be written for the process. Keep in mind that the solid silver chloride does not have a variable concentration, and so is not included in the expression.

$$K_{\mathrm{sp}} = \left[\mathrm{Ag}^{+}\right]\left[\mathrm{Cl}^{-}\right]$$

This equilibrium constant is called the **solubility product constant** $\left(K_{\mathrm{sp}}\right)$ and is equal to the mathematical product of the ions each raised to the power of the coefficient of the ion in the dissociation equation.

The stoichiometry of the formula of the ionic compound dictates the form of the K_{sp} expression. For example, the formula of calcium phosphate is $\mathrm{Ca}_3(\mathrm{PO}_4)_2$. The dissociation equation and K_{sp} expression are shown below:

$$\mathrm{Ca}_3(\mathrm{PO}_4)_2\,(s) \rightleftharpoons 3\mathrm{Ca}^{2+}\,(aq) + 2\mathrm{PO}_4^{3-}\,(aq) \quad K_{\mathrm{sp}} = \left[\mathrm{Ca}^{2+}\right]^3\left[\mathrm{PO}_4^{3-}\right]^2$$

The table below lists solubility product constants for some common nearly insoluble ionic compounds.

Solubility Product Constants $(25^{\circ}\mathrm{C})$

Table 10.18.1: Solubility Product Constants $(25^{\circ}\mathrm{C})$

Compound	K_{sp}	**Compound**	K_{sp}
AgBr	5.0×10^{-13}	CuS	8.0×10^{-37}
AgCl	1.8×10^{-10}	$\mathrm{Fe(OH)}_2$	7.9×10^{-16}
$\mathrm{Al(OH)}_3$	3.0×10^{-34}	$\mathrm{Mg(OH)}_2$	7.1×10^{-12}
BaCO_3	5.0×10^{-9}	PbCl_2	1.7×10^{-5}
BaSO_4	1.1×10^{-10}	PbCO_3	7.4×10^{-14}
CaCO_3	4.5×10^{-9}	PbI_2	7.1×10^{-9}
$\mathrm{Ca(OH)}_2$	6.5×10^{-6}	PbSO_4	6.3×10^{-7}
$\mathrm{Ca}_3(\mathrm{PO}_4)_2$	1.2×10^{-26}	$\mathrm{Zn(OH)}_2$	3.0×10^{-16}
CaSO_4	2.4×10^{-5}	ZnS	3.0×10^{-23}

Summary

- The solubility product constant is equal to the mathematical product of ions each raised to the power of their coefficients in a dissociation equation.
- Calculations using solubility product constants are illustrated.

10.19: Conversion of Solubility to K sp Ksp

Baking soda (sodium bicarbonate) is prepared by bubbling carbon dioxide gas through a solution of ammonia and sodium chloride. Ammonium carbonate is first formed, which then reacts with the NaCl to form sodium bicarbonate and ammonium chloride. The sodium bicarbonate is less soluble than the other materials, so it will precipitate out of solution.

Conversion of Solubility to K_{sp}

Solubility is normally expressed in $\mathrm{g/L}$ of saturated solution. However, solubility can also be expressed in moles per liter. Molar solubility is the number of moles of solute in one liter of saturated solution. In other words, the molar solubility of a given compound represents the highest molarity solution that is possible for that compound. The molar mass of a compound is the conversion factor between solubility and molar solubility. Given that the solubility of $\mathrm{Zn(OH)_2}$ is $4.2 \times 10^{-4}\ \mathrm{g/L}$, the molar solubility can be calculated as shown below:

$$\frac{4.2 \times 10^{-4}\ \cancel{\mathrm{g}}}{\mathrm{L}} \times \frac{1\ \mathrm{mol}}{99.41\ \cancel{\mathrm{g}}} = 4.2 \times 10^{-6}\ \mathrm{mol/L}\ (\mathrm{M})$$

Solubility data can be used to calculate the K_{sp} for a given compound. The following steps need to be taken.

1. Convert from solubility to molar solubility.
2. Use the dissociation equation to determine the concentration of each of the ions in $\mathrm{mol/L}$.
3. Apply the K_{sp} equation.

Example 10.19.1

The solubility of lead (II) fluoride is found experimentally to be $0.533\ \mathrm{g/L}$ Calculate the K_{sp} for lead (II) fluoride.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Solubility of $\mathrm{PbF_2} = 0.533\ \mathrm{g/L}$
- Molar mass $= 245.20\ \mathrm{g/mol}$

Unknown

The dissociation equation for $\mathrm{PbF_2}$ and the corresponding K_{sp} expression:

$$\mathrm{PbF_2}\,(s) \rightleftharpoons \mathrm{Pb^{2+}}\,(aq) + 2\mathrm{F^-}\,(aq) \quad K_{\text{sp}} = \left[\mathrm{Pb^{2+}}\right]\left[\mathrm{F^-}\right]^2$$

The steps above will be followed to calculate the K_{sp} for $\mathrm{PbF_2}$.

Step 2: Solve.

$$\text{molar solubility} = \frac{0.533\ \cancel{\mathrm{g}}}{\mathrm{L}} \times \frac{1\ \mathrm{mol}}{245.20\ \cancel{\mathrm{g}}} = 2.17 \times 10^{-3}\ \mathrm{M}$$

The dissociation equation shows that for every mole of $\mathrm{PbF_2}$ that dissociates, $1\ \mathrm{mol}$ of $\mathrm{Pb^{2+}}$ and $2\ \mathrm{mol}$ of $\mathrm{F^-}$ are produced. Therefore, at equilibrium, the concentrations of the ions are:

$$\left[\mathrm{Pb^{2+}}\right] = 2.17 \times 10^{-3}\ \mathrm{M} \text{ and } \left[\mathrm{F^-}\right] = 2 \times 2.17 \times 10^{-3} = 4.35 \times 10^{-3}\ \mathrm{M}$$

Substitute into the expression and solve for the K_sp.

$$K_\text{sp} = \left(2.17 \times 10^{-3}\right)\left(4.35 \times 10^{-3}\right)^2 = 4.11 \times 10^{-8}$$

Step 3: Think about your result.

The solubility product constant is significantly less than 1 for a nearly insoluble compound such as PbF_2.

Summary

- Molar solubility calculations are described.
- Calculations of K_sp using molar solubility are described.

10.20: Conversion of K sp Ksp to Solubility

Purification of water for drinking and other uses is a complicated process. Heavy metals need to be removed, a process accomplished by the addition of carbonates and sulfates. Lead contamination can present major health problems, especially for children. Lead sulfates and carbonates are very insoluble, so will precipitate out of solution very easily.

Table 10.20.1: Solubility Product Constants (25°C)

Compound	$\boldsymbol{K}_{\text{sp}}$	**Compound**	$\boldsymbol{K}_{\text{sp}}$
AgBr	5.0×10^{-13}	CuS	8.0×10^{-37}
AgCl	1.8×10^{-10}	Fe(OH)_2	7.9×10^{-16}
Al(OH)_3	3.0×10^{-34}	Mg(OH)_2	7.1×10^{-12}
BaCO_3	5.0×10^{-9}	PbCl_2	1.7×10^{-5}
BaSO_4	1.1×10^{-10}	PbCO_3	7.4×10^{-14}
CaCO_3	4.5×10^{-9}	PbI_2	7.1×10^{-9}
Ca(OH)_2	6.5×10^{-6}	PbSO_4	6.3×10^{-7}
$\text{Ca}_3(\text{PO}_4)_2$	1.2×10^{-26}	Zn(OH)_2	3.0×10^{-16}
CaSO_4	2.4×10^{-5}	ZnS	3.0×10^{-23}

The known K_{sp} values from the table above can be used to calculate the solubility of a given compound by following the steps listed below.

1. Set up an ICE problem (Initial, Change, Equilibrium) in order to use the K_{sp} value to calculate the concentration of each of the ions.
2. The concentration of the ions leads to the molar solubility of the compound.
3. Use the molar mass to convert from molar solubility to solubility.

The K_{sp} of calcium carbonate is 4.5×10^{-9}. We begin by setting up an ICE table showing the dissociation of CaCO_3 into calcium ions and carbonate ions. The variable s will be used to represent the molar solubility of CaCO_3. In this case, each formula unit of CaCO_3 yields one Ca^{2+} ion and one CO_3^{2-} ion. Therefore, the equilibrium concentrations of each ion are equal to s.

$$\begin{array}{r|ccccc} & \text{CaCO}_3\,(s) & \rightleftharpoons & \text{Ca}^{2+}\,(aq) & + & \text{CO}_3^{2-}\,(aq) \\ \hline \text{Initial (M)} & & & 0.00 & & 0.00 \\ \text{Change (M)} & & & +s & & +s \\ \text{Equilibrium (M)} & & & s & & s \end{array}$$

The K_{sp} expression can be written in terms of s and then used to solve for s.

$$\begin{aligned} K_{\text{sp}} &= \left[\text{Ca}^{2+}\right]\left[\text{CO}_3^{2-}\right] = (s)(s) = s^2 \\ s &= \sqrt{K_{\text{sp}}} = \sqrt{4.5 \times 10^{-9}} = 6.7 \times 10^{-5}\ \text{M} \end{aligned}$$

The concentration of each of the ions at equilibrium is 6.7×10^{-5} M. We can use the molar mass to convert from molar solubility to solubility.

$$\frac{6.7 \times 10^{-5}\ \cancel{\text{mol}}}{\text{L}} \times \frac{100.09\ \text{g}}{1\ \cancel{\text{mol}}} = 6.7 \times 10^{-3}\ \text{g/L}$$

So, the maximum amount of calcium carbonate that is capable of dissolving in 1 liter of water at 25°C is 6.7×10^{-3} grams. Note that in the case above, the 1:1 ratio of the ions upon dissociation led to the K_{sp} being equal to s^2. This is referred to as a formula of the type AB, where A is the cation and B is the anion. Now let's consider a formula of the type AB_2, such as Fe(OH)_2. In this case, the setup of the ICE table would look like:

	$Fe(OH)_2\,(s)$	$\rightleftharpoons$	$Fe^{2+}\,(aq)$	$+$	$2OH^-\,(aq)$
Initial (M)			0.00		0.00
Change (M)			$+s$		$+2s$
Equilibrium (M)			s		$2s$

When the K_{sp} expression is written in terms of s, we get the following result for the molar solubility.

$$K_{sp} = [Fe^{2+}][OH^-]^2 = (s)(2s)^2 = 4s^3$$

$$s = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{7.9 \times 10^{-16}}{4}} = 5.8 \times 10^{-6}\ M$$

The table below shows the relationship between K_{sp} and molar solubility based on the formula.

Table 10.20.2

Compound Type	Example	K_{sp} Expression	Cation	Anion	K_{sp} in Terms of s
AB	CuS	$[Cu^{2+}][S^{2-}]$	s	s	s^2
AB_2 or A_2B	Ag_2CrO_4	$[Ag^+]^2[CrO_4^{2-}]$	$2s$	s	$4s^3$
AB_3 or A_3B	$Al(OH)_3$	$[Al^{3+}][OH^-]^3$	s	$3s$	$27s^4$
A_2B_3 or A_3B_2	$Ba_3(PO_4)_2$	$[Ba^{2+}]^3[PO_4^{3-}]^2$	$3s$	$2s$	$\mathbf{108s^5}$

The K_{sp} expressions in terms of s can be used to solve problems in which the K_{sp} is used to calculate the molar solubility, as in the examples above. Molar solubility can then be converted to solubility.

Summary

- The process of determining solubility using K_{sp} values from the given table is:
  - Set up an ICE problem and use the K_{sp} value to calculate the concentration of each of the ions.
  - The concentration of the ions leads to the molar solubility of the compound.
  - Use the molar mass to convert from molar solubility to solubility.

- **19.13: Conversion of \(K_\text{sp}\) to Solubility** by CK-12 Foundation is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.21: Predicting Precipitates

The invention of the x-ray machine radically improved medical diagnosis and treatment. For the first time, it was possible to see inside a person's body to detect broken bones, tumors, obstructions, and other problems. Barium sulfate is often used to examine patients with problems of the esophagus, stomach, and intestines. This insoluble compound coats the inside of the tissues and absorbs x-rays, allowing a clear picture of the interior structure of these organs.

Predicting Precipitates

Knowledge of K_{sp} values will allow you to be able to predict whether or not a precipitate will form when two solutions are mixed together. For example, suppose that a known solution of barium chloride is mixed with a known solution of sodium sulfate. Barium sulfate is a mostly insoluble compound, and so could potentially precipitate from the mixture. However, it is first necessary to calculate the **ion product**, $\left[\mathrm{Ba}^{2+}\right]\left[\mathrm{SO}_4^{2-}\right]$, for the solution. If the value of the ion product is less than the value of the K_{sp}, then the solution will remain unsaturated; no precipitate will form because the concentrations are not high enough to begin the precipitation process. If the value of the ion product is greater than the value of the K_{sp}, then a precipitate will form. The formation of the precipitate lowers the concentration of each of the ions until the ion product is exactly equal to the K_{sp}, at which point precipitation ceases.

Figure 10.21.1: Barium sulfate is used as a component of white pigment for paints and as an agent in certain x-ray imaging processes.

Example 10.21.1

Will a precipitate of barium sulfate form when $10.0\ \mathrm{mL}$ of $0.0050\ \mathrm{M}\ \mathrm{BaCl_2}$ is mixed with $20.0\ \mathrm{mL}$ of $0.0020\ \mathrm{M}\ \mathrm{Na_2SO_4}$?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Concentration of $\mathrm{BaCl_2} = 0.0050\ \mathrm{M}$
- Volume of $\mathrm{BaCl_2} = 10.0\ \mathrm{mL}$
- Concentration of $\mathrm{Na_2SO_4} = 0.0020\ \mathrm{M}$
- Volume of $\mathrm{Na_2SO_4} = 20.0\ \mathrm{mL}$
- K_{sp} of $\mathrm{BaSO_4} = 1.1 \times 10^{-10}$

Unknown

- Ion product $= \left[\mathrm{Ba}^{2+}\right]\left[\mathrm{SO}_4^{2-}\right]$
- If a precipitate forms.

The concentration and volume of each solution that is mixed together must be used to calculate the $\left[\mathrm{Ba}^{2+}\right]$ and the $\left[\mathrm{SO}_4^{2-}\right]$. Each individual solution is diluted when they are mixed together. The ion product is calculated and compared to the K_{sp} to determine if a precipitate forms.

Step 2: Solve.

The moles of each ion from the original solutions are calculated by multiplying the molarity by the volume in liters.

$$\text{mol Ba}^{2+} = 0.0050 \text{ M} \times 0.010 \text{ L} = 5.0 \times 10^{-5} \text{ mol Ba}^{2+}$$

$$\text{mol SO}_4^{2-} = 0.0020 \text{ M} \times 0.020 \text{ L} = 4.0 \times 10^{-5} \text{ mol SO}_4^{2-}$$

The concentration of each ion after dilution is then calculated by dividing the moles by the final solution volume of 0.030 L

$$\left[\text{Ba}^{2+}\right] = \frac{5.0 \times 10^{-5} \text{ mol}}{0.030 \text{ L}} = 1.7 \times 10^{-3} \text{ M}$$

$$\left[\text{SO}_4^{2-}\right] = \frac{4.0 \times 10^{-5} \text{ mol}}{0.030 \text{ L}} = 1.3 \times 10^{-3} \text{ M}$$

Now, the ion product is calculated.

$$\left[\text{Ba}^{2+}\right]\left[\text{SO}_4^{2-}\right] = \left(1.7 \times 10^{-3}\right)\left(1.3 \times 10^{-3}\right) = 2.2 \times 10^{-6}$$

Since the ion product is greater than the K_{sp}, a precipitate of barium sulfate will form.

Step 3: Think about your result.

Two significant figures are appropriate for the calculated value of the ion product.

Summary

10.22: Common Ion Effect

Lithium carbonate is an essential component of lithium batteries, which tend to be longer-lasting than regular alkaline batteries. The material is obtained from lithium ores by adding $\mathrm{CO_2}$ under high pressure to form the more soluble $\mathrm{LiHCO_3}$. The mixture is then de-pressurized to remove the carbon dioxide, and the lithium carbonate precipitates out of solution.

Common Ion Effect

In a saturated solution of calcium sulfate, an equilibrium exists between the solid calcium sulfate and its ions in solution:

$$\mathrm{CaSO_4}\,(s) \rightleftharpoons \mathrm{Ca^{2+}}\,(aq) + \mathrm{SO_4^{2-}}\,(aq) \quad K_{\mathrm{sp}} = 2.4 \times 10^{-5}$$

Suppose that some calcium nitrate was added to this saturated solution. Immediately, the concentration of the calcium ion in the solution would increase. As a result, the ion product of the $\left[\mathrm{Ca^{2+}}\right]$ times the $\left[\mathrm{SO_4^{2-}}\right]$ would increase and now be greater than the K_{sp}. According to Le Chatelier's principle, the equilibrium above would shift to the left in order to relieve the stress of the added calcium ion. Additional calcium sulfate would precipitate out of the solution until the ion product once again becomes equal to the K_{sp}. Note that in the new equilibrium, the concentrations of the calcium ion and the sulfate ion would no longer be equal to each other. The calcium ion concentration would be larger than the sulfate ion concentration.

This situation describes the common ion effect. A **common ion** is an ion that is in common to both salts in a solution. In the above example, the common ion is $\mathrm{Ca^{2+}}$. The **common ion effect** is a decrease in the solubility of an ionic compound as a result of the addition of a common ion. Adding calcium ion to the saturated solution of calcium sulfate causes additional $\mathrm{CaSO_4}$ to precipitate from the solution, lowering its solubility. The addition of a solution containing sulfate ion, such as potassium sulfate, would result in the same common ion effect.

Example 10.22.1

What is the concentration of zinc ion in $1.00\ \mathrm{L}$ of a saturated solution of zinc hydroxide to which $0.040\ \mathrm{mol}$ of NaOH has been added?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $K_{\mathrm{sp}} = 3.0 \times 10^{-16}$ (from table in "Conversion of K_{sp} to Solubility")
- Moles of added $\mathrm{NaOH} = 0.040\ \mathrm{mol}$
- Volume of solution $= 1.00\ \mathrm{L}$

Unknown

Express the concentrations of the two ions relative to the variable s. The concentration of the zinc ion will be equal to s, while the concentration of the hydroxide ion will be equal to $0.040 + 2s$.

Step 2: Solve.

The K_{sp} expression can be written in terms of the variable s.

$$K_{\mathrm{sp}} = \left[\mathrm{Zn^{2+}}\right]\left[\mathrm{OH^-}\right]^2 = (s)\,(0.040 + 2s)^2$$

Because the value of the K_{sp} is so small, we can make the assumption that the value of s will be very small compared to 0.040. This simplifies the mathematics involved in solving for s.

$$\begin{aligned} K_{\text{sp}} &= (s)\,(0.040)^2 = 0.0016s = 3.0 \times 10^{-16} \\ s &= \frac{K_{\text{sp}}}{\left[\text{OH}^-\right]^2} = \frac{3.0 \times 10^{-16}}{0.0016} = 1.9 \times 10^{-13}\ \text{M} \end{aligned}$$

The concentration of the zinc ion is equal to s and so $\left[\text{Zn}^{2+}\right] = 1.9 \times 10^{-13}\ \text{M}$.

Step 3: Think about your result.

The relatively high concentration of the common ion, OH^-, results in a very low concentration of zinc ion. The molar solubility of the zinc hydroxide is less in the presence of the common ion than it would be in water.

Summary

- The common ion is an ion that is in common to both salts in a solution.
- The common ion effect is a decrease in the solubility of an ionic compound as a result of the addition of a common ion.
- Calculations involving the common ion effect are described.

CHAPTER OVERVIEW

11: Acids and Bases

11.1: Properties of Acids

A cup of coffee first thing in the morning can help start the day. However, keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee!

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food.

Acids are a distinct class of compounds because of the properties of their aqueous solutions. These properties are:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g)$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

Summary

11.2: Properties of Bases

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.

Figure* $11.2.1$*:* *Phenolphthalein indicator in presence of base.

Please note that tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words, don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate.

Summary

11.3: Arrhenius Acids

Venus is the planet nearest to Earth, but has a very different and hostile environment. It has a surface temperature that averages around 450^oC. The atmosphere is composed of carbon dioxide, but clouds of sulfuric acid move through the upper atmosphere, helping to create the extremely unfriendly conditions. Because of these conditions, Venus is not a place you want to visit on vacation.

Arrhenius Acids

Swedish chemist Svante Arrhenius (1859-1927) was the first to propose a theory to explain the observed behavior of acids and bases. Because of their ability to conduct a current, he knew that both acids and bases contained ions in solution. An **Arrhenius acid** is a compound that ionizes to yield hydrogen ions $\left(\text{H}^+ \right)$ in aqueous solution.

Acids are molecular compounds with ionizable hydrogen atoms. Only hydrogen atoms that are part of a highly polar covalent bond are ionizable. Hydrogen chloride $\left(\text{HCl} \right)$ is a gas at room temperature and under normal pressure. The $\text{H}-\text{Cl}$ bond in hydrogen chloride is a polar bond. The hydrogen atom is electron deficient because of the higher electronegativity of the chlorine atom. Consequently, the hydrogen atom is attracted to the lone pair of electrons in a water molecule when HCl is dissolved in water. The result is that the $\text{H}-\text{Cl}$ bond breaks, with both bonding electrons remaining with the Cl, forming a chloride ion. The H^+ ion attaches to the water molecule, forming a polyatomic ion called the hydronium ion. The **hydronium ion** $\left(\text{H}_3\text{O}^+ \right)$ can be thought of as a water molecule with an extra attached hydrogen ion.

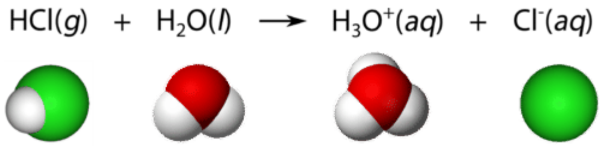

Figure 11.3.1: Formation of a hydronium ion. (CC BY-NC; CK-12)

Equations showing the ionization of an acid in water are frequently simplified by omitting the water molecule:

$$\text{HCl} \left(g \right) \rightarrow \text{H}^+ \left(aq \right) + \text{Cl}^- \left(aq \right)$$

This is merely a simplification of the previous equation, but it is commonly used. Any hydrogen ions in an aqueous solution will be attached to water molecules as hydronium ions.

Not all hydrogen atoms in molecular compounds are ionizable. In methane $\left(\text{CH}_4 \right)$, the hydrogen atoms are covalently bonded to carbon in bonds that are only slightly polar. The hydrogen atoms are not capable of ionizing, and methane has no acidic properties. Acetic acid $\left(\text{CH}_3\text{COOH} \right)$ belongs to a class of acids called organic acids. There are four hydrogen atoms in the molecule, but only the one hydrogen that is bonded to an oxygen atom is ionizable.

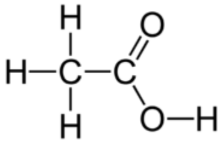

Figure 11.3.2: The $\text{O}-\text{H}$ bond can be ionized to yield the H^+ ion and the acetate ion. The other hydrogen atoms in this molecule are not acidic. (CC BY-NC; CK-12)

The table below lists some of the more common acids:

Common Acids

Table 11.3.1: Common Acids

Acid Name	Formula
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Acetic acid	CH_3COOH

Table 11.3.1: Common Acids

Hypochlorous acid	$HClO$

A **monoprotic acid** is an acid that contains only one ionizable hydrogen. Hydrochloric acid and acetic acid are monoprotic acids. A **polyprotic acid** is an acid that contains multiple ionizable hydrogens. Most common polyprotic acids are either diprotic (such as H_2SO_4) or triprotic (such as H_3PO_4).

Summary

- An Arrhenius acid is a compound that ionizes to yield hydrogen ions $\left(H^+\right)$ in aqueous solution.
- Examples of Arrhenius acids are given.

11.4: Arrhenius Bases

Sodium hydroxide is a versatile chemical. It can be used for such mundane purposes as cleaning clogged drains. Several commercial preparations contain sodium hydroxide for this purpose. It also has a number of applications in the food processing field. Ice cream is thickened using $NaOH$. If olives are soaked in a solution containing sodium hydroxide and other chemicals, the olives will turn black. Soft pretzels are made chewy by the application of sodium hydroxide to the food. This compound has been widely used in the synthesis of plastics, for etching aluminum, for paint removal, and is employed in the dehorning of cattle.

Arrhenius Bases

An **Arrhenius base** is a compound that ionizes to yield hydroxide ions $\left(OH^-\right)$ in aqueous solution. The table below lists several of the more common bases.

Table 11.4.1: Common Bases

Base Name	Formula
Sodium hydroxide	$NaOH$
Potassium hydroxide	KOH
Magnesium hydroxide	$Mg(OH)_2$
Calcium hydroxide	$Ca(OH)_2$

All of the bases listed in the table are solids at room temperature. Upon dissolving in water, each dissociates into a metal cation and the hydroxide ion.

$$NaOH\left(s\right) \rightarrow Na^+\left(aq\right) + OH^-\left(aq\right)$$

Sodium hydroxide is a very caustic substance also known as lye. Lye is used as a rigorous cleaner and is an ingredient in the manufacture of soaps. Care must be taken with strong bases like sodium hydroxide, as exposure can lead to severe burns (see figure below).

Figure 11.4.1: This foot has severe burns due to prolonged contact with a solution of sodium hydroxide, also known as lye. (CC BY-NC; CK-12)

Sodium belongs to the group of elements called the alkali metals. An **alkaline solution** is another name for a solution that is basic. All alkali metals react readily with water to produce the metal hydroxide and hydrogen gas. The resulting solutions are basic.

$$2K\left(s\right) + 2H_2O\left(l\right) \rightarrow 2KOH\left(aq\right) + H_2\left(g\right)$$

Bases that consist of an alkali metal cation and the hydroxide anion are all very soluble in water. Compounds of the Group 2 metals (the alkaline earth metals) are also basic. However, these compounds are generally not as soluble in water. Therefore, the dissociation reactions for these compounds are shown as equilibrium reactions:

$$Mg(OH)_2\left(s\right) \overset{H_2O}{\rightleftharpoons} Mg^{2+}\left(aq\right) + 2OH^-\left(aq\right)$$

These relatively insoluble hydroxides were some of the compounds discussed in the context of the solubility product constant $\left(K_{sp}\right)$. The solubility of magnesium hydroxide is $0.0084\,g$ per liter of water at $25^\circ C$. Because of its low solubility, magnesium hydroxide is not as dangerous as sodium hydroxide. In fact, magnesium hydroxide is the active ingredient in a product called milk of magnesia, which is used as an antacid or a mild laxative.

Summary

- Arrhenius base is a compound that ionizes to yield hydroxide ions $\left(\mathrm{OH}^{-}\right)$ in aqueous solution.
- Examples of Arrhenius bases are given.

11.5: Brønsted-Lowry Acids and Bases

The Arrhenius concept of acids and bases was a significant contribution to the scientific understanding of acids and bases. It replaced and expanded Lavoisier's original idea that all acids contained oxygen. However, the Arrhenius theory also had its shortcomings. It did not take into account the role of the solvent. In addition, this concept did not deal with acid-base behavior in solvents such as benzene, where there could be no ionization. So, formation of a new theory was imperative, which built on the findings of Arrhenius, but also went beyond them.

Brønsted-Lowry Acids and Bases

The Arrhenius definition of acids and bases is somewhat limited. There are some compounds whose properties suggest that they are either acidic or basic, but which do not qualify according to the Arrhenius definition. An example is ammonia $\left(\mathrm{NH_3}\right)$. Its aqueous solution turns litmus blue, it reacts with acids, and displays all the other properties of a base. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by Danish chemist Johannes Brønsted (1879-1947) and English chemist Thomas Lowry (1874-1936). A **Brønsted-Lowry acid** is a molecule or ion that donates a hydrogen ion in a reaction. A **Brønsted-Lowry base** is a molecule or ion that accepts a hydrogen ion in a reaction. A hydrogen ion is commonly referred to as a proton, and so acids and bases are proton donors and proton acceptors, respectively, according to the Brønsted-Lowry definition. All substances that are categorized as acids and bases under the Arrhenius definition are also defined as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

What kind of molecule would qualify as a Brønsted-Lowry base? These molecules need to be able to accept a hydrogen ion (or proton). Two possibilities come to mind: an anion that can form a neutral compound with a proton, or a molecule in which one or more atoms has lone-pair electrons. The most obvious anion is the Arrhenius base $\mathrm{OH^-}$. This ion can form a water molecule with a proton by accepting the proton. The acetate anion $\mathrm{CH_3COO^-}$ is another anion that can combine with a proton to form acetic acid. Lone-pair electron groups would include the nitrogen atom (see figure below). The two electrons at the top of the nitrogen atom are not connected in any type of bond, but they do interact readily with a bare proton.

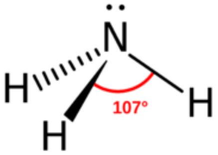

Figure 11.5.1: *Lone pair electrons on nitrogen.* (CC BY-NC; CK-12)

Oxygen is another atom with lone pair electrons that can function as a Brønsted-Lowry base.

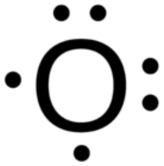

Figure 11.5.2: Electrons in an oxygen atom. (CC BY-NC; CK-12)

The two single electrons (to the left and bottom of the atom) can form single covalent bonds with other atoms, while the two pairs of double electrons (top and right) are available to interact with a hydrogen ion.

Summary

- A Brønsted-Lowry acid is a molecule or ion that donates a hydrogen ion in a reaction.
- A Brønsted-Lowry base is a molecule or ion that accepts a hydrogen ion in a reaction.

11.6: Brønsted-Lowry Acid-Base Reactions

The Roman god Janus was considered the god of gates, doors, beginnings, and endings. He is portrayed with two faces, looking in two directions at once. Janus would have been in a good position to look at the acid-base reactions we see in this concept, since they are equilibrium reactions involving two different forms of both acids and bases.

Brønsted-Lowry Acid-Base Reactions

An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another. When ammonia is dissolved in water, it undergoes the following reversible reaction.

$$\underset{\text{base}}{\mathrm{NH_3}\,(aq)} + \underset{\text{acid}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{NH_4^+}\,(aq)} + \underset{\text{base}}{\mathrm{OH^-}\,(aq)}$$

In this reaction, the water molecule is donating a proton to the ammonia molecule. The resulting products are the ammonium ion and the hydroxide ion. The water is acting as a Brønsted-Lowry acid, while the ammonia is acting as a Brønsted-Lowry base. The hydroxide ion that is produced causes the solution to be basic.

We can also consider the reverse reaction in the above equation. In that reaction, the ammonia ion donates a proton to the hydroxide ion. The ammonium ion is a Brønsted-Lowry acid, while the hydroxide ion is a Brønsted-Lowry base. Most Brønsted-Lowry acid-base reactions can be analyzed in this way. One acid and one base are reactants, and one acid and one base are products.

In the above reaction, water acted as an acid, which may seem a bit unexpected. Water can also act as a base in a Brønsted-Lowry acid-base reaction, as long as it reacts with a substance that is a better proton donor. Shown below is the reaction of water with the hydrogen sulfate ion.

$$\underset{\text{acid}}{\mathrm{HSO_4^-}\,(aq)} + \underset{\text{base}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{H_3O^+}\,(aq)} + \underset{\text{base}}{\mathrm{SO_4^{2-}}\,(aq)}$$

So, water is capable of being either an acid or a base, a characteristic called amphoterism. An **amphoteric substance** is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.

Conjugate Acids and Bases

When a substance that is acting as a Brønsted-Lowry acid donates its proton, it becomes a base in the reverse reaction. In the reaction above, the hydrogen sulfate ion $\left(\mathrm{HSO_4^-}\right)$ donates a proton to water and becomes a sulfate ion $\left(\mathrm{SO_4^{2-}}\right)$. The $\mathrm{HSO_4^-}$ and the $\mathrm{SO_4^{2-}}$ are linked to one another by the presence or absence of the $\mathrm{H^+}$ ion. A **conjugate acid-base pair** is a pair of substances related by the loss or gain of a single hydrogen ion. A **conjugate acid** is the particle produced when a base accepts a proton. The hydrogen sulfate ion is the conjugate acid of the sulfate ion. A **conjugate base** is the particle produced when an acid donates a proton. The sulfate ion is the conjugate base of the hydrogen sulfate ion.

In the reaction illustrated below, water serves both as acid and base simultaneously. One water molecule serves as an acid and donates a proton. The other water molecule functions as a base by accepting the proton.

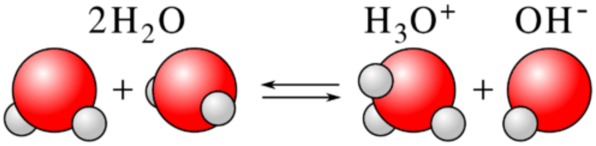

Figure 11.6.1: Water molecules as a conjugate acid-base pair. (CC BY-NC; CK-12)

A typical Brønsted-Lowry acid-base reaction contains two conjugate acid-base pairs, as shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{PO_4^{3-}}\,(aq) \rightleftharpoons \mathrm{NO_2^-}\,(aq) + \mathrm{HPO_4^{2-}}\,(aq)$$

One conjugate acid-base pair is $\mathrm{NHO_2/NO_2^-}$, while the other pair is $\mathrm{HPO_4^{2-}/PO_4^{3-}}$.

Summary

- An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another.
- A conjugate acid-base pair is a pair of substances related by the loss or gain of a single hydrogen ion.
- A conjugate acid is the particle produced when a base accepts a proton.
- A conjugate base is the particle produced when an acid donates a proton.
- Examples of conjugate acid-base pairs are given.

11.7: Lewis Acids and Bases

Ideas in science do not stay static. One discovery builds upon another. The concept of acids and bases has grown from the fundamental ideas of Arrhenius to Brønsted-Lowry to Lewis. Each step adds to our understanding of the surrounding world, and makes the "big picture" even bigger.

Lewis Acids and Bases

Gilbert Lewis (1875-1946) proposed a third theory of acids and bases that is even more general than either the Arrhenius or Brønsted-Lowry theories. A **Lewis acid** is a substance that accepts a pair of electrons to form a covalent bond. A **Lewis base** is a substance that donates a pair of electrons to form a covalent bond. So, a Lewis acid-base reaction is represented by the transfer of a pair of electrons from a base to an acid. A hydrogen ion, which lacks any electrons, accepts a pair of electrons. It is an acid under both the Brønsted-Lowry and Lewis definitions. Ammonia consists of a nitrogen atom as the central atom with a lone pair of electrons. The reaction between ammonia and the hydrogen ion can be depicted as shown in the figure below.

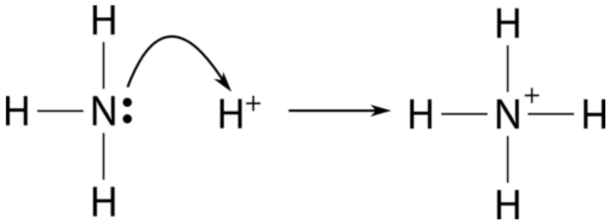

Figure 11.7.1: Reaction between ammonia and a proton. (CC BY-NC; CK-12)

The lone pair on the nitrogen atom is transferred to the hydrogen ion, making the NH_3 a Lewis base while the H^+ is a Lewis acid.

Some reactions that do not qualify as acid-base reactions under the other definitions do so under only the Lewis definition. An example is the reaction of ammonia with boron trifluoride.

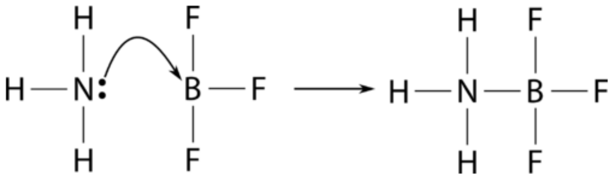

Figure 11.7.2: Ammonia and boron trifluoride. (CC BY-NC; CK-12)

Boron trifluoride is the Lewis acid, while ammonia is again the Lewis base. As there is no hydrogen ion involved in this reaction, it qualifies as an acid-base reaction only under the Lewis definition. The table below summarizes the three acid-base theories.

Acid-Base Definitions

Table 11.7.1: Acid-Base Definitions

Type	Acid	Base
Arrhenius	H^+ ions in solution	OH^- ions in solution
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

Summary

- A Lewis acid is a substance that accepts a pair of electrons to form a covalent bond.
- A Lewis base is a substance that donates a pair of electrons to form a covalent bond.
- Examples of Lewis acids and bases are given.

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11.8: Ion-Product of Water

At one time, you could take the little caps off the top of a car battery and check the condition of the sulfuric acid inside. If it got low, you could add more acid. But, sulfuric acid is hazardous stuff, so the batteries are now sealed to protect people. Because of the acid's dangerous nature, it is not a good idea to cut into a battery to see what it looks like—you could get acid burns.

The Ion-Product of Water

The **self-ionization of water** (the process in which water ionizes to hydronium ions and hydroxide ions) occurs to a very limited extent. When two molecules of water collide, there can be a transfer of a hydrogen ion from one molecule to the other. The products are a positively charged hydronium ion and a negatively charged hydroxide ion.

$$\mathrm{H_2O}\,(l) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

We often use the simplified form of the reaction:

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The equilibrium constant for the self-ionization of water is referred to as the ion-product for water and is given the symbol K_w.

$$K_\mathrm{w} = \left[\mathrm{H^+}\right]\left[\mathrm{OH^-}\right]$$

The **ion-product of water** $(\boldsymbol{K}_\mathrm{w})$ is the mathematical product of the concentration of hydrogen ions and hydroxide ions. Note that $\mathrm{H_2O}$ is not included in the ion-product expression because it is a pure liquid. The value of K_w is very small, in accordance with a reaction that favors the reactants. At $25^\mathrm{o}\mathrm{C}$, the experimentally determined value of K_w in pure water is 1.0×10^{-14}.

$$K_\mathrm{w} = \left[\mathrm{H^+}\right]\left[\mathrm{OH^-}\right] = 1.0 \times 10^{-14}$$

In pure water, the concentrations of hydrogen and hydroxide ions are equal to one another. Pure water or any other aqueous solution in which this ratio holds is said to be neutral. To find the molarity of each ion, the square root of K_w is taken.

$$\left[\mathrm{H^+}\right] = \left[\mathrm{OH^-}\right] = 1.0 \times 10^{-7}$$

An **acidic solution** is a solution in which the concentration of hydrogen ions is greater than the concentration of hydroxide ions. For example, hydrogen chloride ionizes to produce $\mathrm{H^+}$ and $\mathrm{Cl^-}$ ions upon dissolving in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

This increases the concentration of $\mathrm{H^+}$ ions in the solution. According to Le Chatelier's principle, the equilibrium represented by $\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$ is forced to the left, towards the reactant. As a result, the concentration of the hydroxide ion decreases.

A **basic solution** is a solution in which the concentration of hydroxide ions is greater than the concentration of hydrogen ions. Solid potassium hydroxide dissociates in water to yield potassium ions and hydroxide ions.

$$\mathrm{KOH}\,(s) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The increase in concentration of the $\mathrm{OH^-}$ ions causes a decrease in the concentration of the $\mathrm{H^+}$ ions and the ion-product of $\left[\mathrm{H^+}\right]\left[\mathrm{OH^-}\right]$ remains constant.

Example 11.8.1

Hydrochloric acid (HCl) is a strong acid, meaning it is 100% ionized in solution. What is the $\left[\mathrm{H^+}\right]$ and the $\left[\mathrm{OH^-}\right]$ in a solution of $2.0 \times 10^{-3}\ \mathrm{M\,HCl}$?

Solution

Step 1: List the known values and plan the problem.

Known

- $[\mathrm{HCl}] = 2.0 \times 10^{-3}\ \mathrm{M}$
- $K_\mathrm{w} = 1.0 \times 10^{-14}$

Unknown

- $\left[\mathrm{H}^+\right] = ?\ \mathrm{M}$
- $\left[\mathrm{OH}^-\right] = ?\ \mathrm{M}$

Because HCl is 100% ionized, the concentration of H^+ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one H^+ ion and one Cl^- ion. The concentration of OH^- can then be determined from the $\left[\mathrm{H}^+\right]$ and K_w.

Step 2: Solve.

$$\begin{aligned} \left[\mathrm{H}^+\right] &= 2.0 \times 10^{-3}\ \mathrm{M} \\ K_\mathrm{w} &= \left[\mathrm{H}^+\right]\left[\mathrm{OH}^-\right] = 1.0 \times 10^{-14} \\ \left[\mathrm{OH}^-\right] &= \frac{K_\mathrm{w}}{\left[\mathrm{H}^+\right]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12}\ \mathrm{M} \end{aligned}$$

Step 3: Think about your result.

The $\left[\mathrm{H}^+\right]$ is much higher than the $\left[\mathrm{OH}^-\right]$ because the solution is acidic. As with other equilibrium constants, the unit for K_w is customarily omitted.

Summary

- The self-ionization of water is described and an ionization constant for the process is stated.
- Acidic and basic solutions are defined.
- Calculations using K_w are illustrated.

11.9: The pH Scale

Grapefruit juice has a pH of somewhere between 2.9 and 3.3, depending on the specific product. Excessive exposure to this juice can cause erosion of tooth enamel and can lead to tooth damage. The acids in grapefruit juice are carbon-based, with citric acid being one of the major constituents. This compound has three ionizable hydrogens on each molecule which contribute to the relatively low pH of the juice. Another component of grapefruit juice is malic acid, containing two ionizable hydrogens per molecule.

The pH Scale

Expressing the acidity of a solution by using the molarity of the hydrogen ion is cumbersome because the quantities are generally very small. Danish scientist Søren Sørensen (1868-1939) proposed an easier system for indicating the concentration of $\mathrm{H^+}$ called the pH scale. The letters pH stand for the power of the hydrogen ion. The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration:

$$\mathrm{pH} = -\log\left[\mathrm{H^+}\right]$$

In pure water or a neutral solution, the $\left[\mathrm{H^+}\right] = 1.0 \times 10^{-7}\ \mathrm{M}$. Substituting into the pH expression:

$$\mathrm{pH} = -\log\left[1.0 \times 10^{-7}\right] = -(-7.00) = 7.00$$

The pH of pure water or any neutral solution is thus 7.00. For recording purposes, the numbers to the right of the decimal point in the pH value are the significant figures. Since 1.0×10^{-7} has two significant figures, the pH can be reported as 7.00.

A logarithmic scale condenses the range of acidity to numbers that are easy to use. Consider a solution with $\left[\mathrm{H^+}\right] = 1.0 \times 10^{-4}\ \mathrm{M}$. That is a hydrogen-ion concentration that is 1000 times higher than the concentration in pure water. The pH of such a solution is 4.00, a difference of just 3 pH units. Notice that when the $\left[\mathrm{H^+}\right]$ is written in scientific notation and the coefficient is 1, the pH is simply the exponent with the sign changed. The pH of a solution with $\left[\mathrm{H^+}\right] = 1 \times 10^{-2}\ \mathrm{M}$ is 2 and the pH of a solution with $\left[\mathrm{H^+}\right] = 1 \times 10^{-10}\ \mathrm{M}$ is 10.

As we saw earlier, a solution with $\left[\mathrm{H^+}\right]$ higher than 1.0×10^{-7} is acidic, while a solution with $\left[\mathrm{H^+}\right]$ lower than 1.0×10^{-7} is basic. Consequently, solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic. Figure 11.9.1 illustrates this relationship, along with some examples of various solutions.

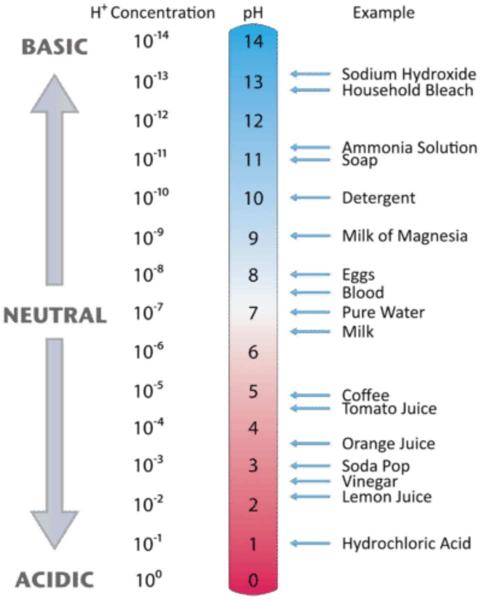

Figure 11.9.1: The pH values for several common materials. (CC BY-NC; CK-12)

Summary

- The pH of a solution is the negative logarithm of the hydrogen-ion concentration.
- Solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic.

- pH values for several common materials are listed.

11.10: Calculating pH of Acids and Bases

It's not uncommon to see a tropical fish tank in homes or businesses. These brightly-colored creatures are relaxing to watch, but do require a certain amount of maintenance in order for them to survive. Tap water is usually too alkaline when it comes out of the faucet, so some adjustments need to be made. The pH of the water will change over time while it is in the tank, which means it needs to be tested every so often. Any fish tank caretaker has a chance to be a chemist for their fish!

Calculating pH of Acids and Bases

Calculation of pH is simple when there is a $1 \times 10^{\text{power}}$ problem. However, in real life, this is rarely the situation. If the coefficient is not equal to 1, a calculator must be used to find the pH. For example, the pH of a solution with $\left[\mathrm{H}^{+}\right]=2.3 \times 10^{-5} \mathrm{~M}$ can be found as shown below.

$$\mathrm{pH}=-\log \left[2.3 \times 10^{-5}\right]=4.64$$

When the pH of a solution is known, the concentration of the hydrogen ion can be calculated. The inverse of the logarithm (or antilog) is the 10^{x} key on a calculator.

$$\left[\mathrm{H}^{+}\right]=10^{-\mathrm{pH}}$$

For example, suppose that you have a solution with a pH of 9.14. To find the $\left[\mathrm{H}^{+}\right]$ use the 10^{x} key.

$$\left[\mathrm{H}^{+}\right]=10^{-\mathrm{pH}}=10^{-9.14}=7.24 \times 10^{-10} \mathrm{~M}$$

Hydroxide Ion Concentration and pH

As we saw earlier, the hydroxide ion concentration of any aqueous solution is related to the hydrogen ion concentration through the value of K_{w}. We can use that relationship to calculate the pH of a solution of a base.

Example 11.10.1

Sodium hydroxide is a strong base. Find the pH of a solution prepared by dissolving $1.0 \mathrm{~g}$ of NaOH into enough water to make $1.0 \mathrm{~L}$ of solution.

Solution

Step 1: List the known values and plan the problem.

Known

- Mass $\mathrm{NaOH}=1.0 \mathrm{~g}$
- Molar mass $\mathrm{NaOH}=40.00 \mathrm{~g} / \mathrm{mol}$
- Volume solution $=1.0 \mathrm{~L}$
- $K_{\mathrm{w}}=1.0 \times 10^{-14}$

Unknown

First, convert the mass of NaOH to moles. Second, calculate the molarity of the NaOH solution. Because NaOH is a strong base and is soluble, the $\left[\mathrm{OH}^{-}\right]$ will be equal to the concentration of the NaOH. Third, use K_{w} to calculate the $\left[\mathrm{H}^{+}\right]$ in the solution. Lastly, calculate the pH.

Step 2: Solve.

$$1.00 \text{ g } \cancel{\text{NaOH}} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g } \cancel{\text{NaOH}}} = 0.025 \text{ mol NaOH}$$

$$\text{Molarity} = \frac{0.025 \text{ mol NaOH}}{1.00 \text{ L}} = 0.025 \text{ M NaOH} = 0.025 \text{ M OH}^-$$

$$\left[\text{H}^+\right] = \frac{K_\text{w}}{\left[\text{OH}^-\right]} = \frac{1.0 \times 10^{-14}}{0.025 \text{ M}} = 4.0 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log\left[\text{H}^+\right] = -\log\left(4.0 \times 10^{-13}\right) = 12.40$$

Step 3: Think about your result.

The solution is basic and so its pH is greater than 7. The reported pH is rounded to two decimal places because the original mass and volume has two significant figures.

Summary

11.11: The pOH Concept

Soap Lake has a long history as a healing place. Indian tribes would put aside their rivalries when they came to the lake to enjoy the high mineral content of the water. In the days before good antibiotics, thousands of visitors would come and enjoy the soothing waters of the lake. Soap Lake is alkaline, with water quality though to be similar to that of the moons of the planet Jupiter.

The pOH Concept

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration:

$$\mathrm{pOH} = -\log\left[\mathrm{OH^-}\right]$$

The pH of a solution can be related to the pOH. Consider a solution with a pH $= 4.0$. The $\left[\mathrm{H^+}\right]$ of the solution would be $1.0 \times 10^{-4}\ \mathrm{M}$. Dividing K_w by this yields a $\left[\mathrm{OH^-}\right]$ of $1.0 \times 10^{-10}\ \mathrm{M}$. Finally the pOH of the solution equals $-\log\left(1.0 \times 10^{-10}\right) = 10$. This example illustrates the following relationship.

$$\mathrm{pH} + \mathrm{pOH} = 14$$

The pOH scale is similar to the pH scale in that a pOH of 7 is indicative of a neutral solution. A basic solution has a pOH less than 7, while an acidic solution has a pOH of greater than 7. The pOH is convenient to use when finding the hydroxide ion concentration from a solution with a known pH.

Example 11.11.1

Find the hydroxide concentration of a solution with a pH of 4.42.

Solution

Step 1: List the known values and plan the problem.

Known

- $\mathrm{pH} = 4.42$
- $\mathrm{pH} + \mathrm{pOH} = 14$

Unknown

First, the pOH is calculated, followed by the $\left[\mathrm{OH^-}\right]$.

Step 2: Solve.

$$\begin{aligned} \mathrm{pOH} &= 14 - \mathrm{pH} = 14 - 4.42 = 9.58 \\ \left[\mathrm{OH^-}\right] &= 10^{-\mathrm{pOH}} = 10^{-9.58} = 2.6 \times 10^{-10}\ \mathrm{M} \end{aligned}$$

Step 3: Think about your result.

The pH is that of an acidic solution, and the resulting hydroxide-ion concentration is less than $1 \times 10^{-7}\ \mathrm{M}$. The answer has two significant figures because the given pH has two decimal places.

The diagram below shows all of the interrelationships between $\left[\mathrm{H^+}\right]$, $\left[\mathrm{OH^-}\right]$, pH, and pOH.

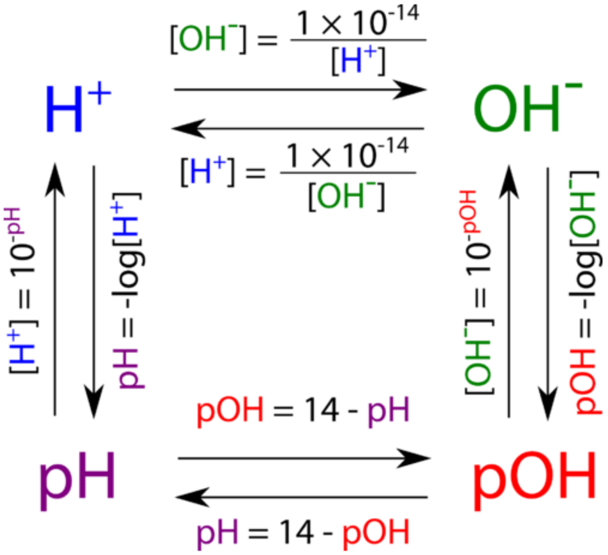

Figure 11.11.1: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH. (CC BY-NC; CK-12)

Summary

- The pOH of a solution is the negative logarithm of the hydroxide-ion concentration.
- Calculations involving pOH are described.

11.12: Strong and Weak Acids and Acid Ionization Constant (K a) (Ka)

The etching of glass is a slow process that can produce beautiful artwork. Traditionally, the glass has been treated with dilute hydrofluoric acid which gradually dissolves the glass under it. Parts of the piece that should not be etched are covered with wax or some other non-reactive material. In more recent times, compounds such as ammonium bifluoride have been used. Whichever chemical is employed, the artist must be very careful not to get any on their skin.

Strong and Weak Acids and Acid Ionization Constant

Acids are classified as either strong or weak, based on their ionization in water. A **strong acid** is an acid which is completely ionized in an aqueous solution. Hydrogen chloride (HCl) ionizes completely into hydrogen ions and chloride ions in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^+\,(aq) + \mathrm{Cl}^-\,(aq)$$

A **weak acid** is an acid that ionizes only slightly in an aqueous solution. Acetic acid (found in vinegar) is a very common weak acid. Its ionization is shown below.

$$\mathrm{CH_3COOH}\,(aq) \rightleftharpoons \mathrm{H}^+\,(aq) + \mathrm{CH_3COO}^-\,(aq)$$

The ionization of acetic acid is incomplete, and so the equation is shown with a double arrow. The extent of ionization of weak acids varies, but is generally less than 10%. A $0.10\ \mathrm{M}$ solution of acetic acid is only about 1.3% ionized, meaning that the equilibrium strongly favors the reactants.

Weak acids, like strong acids, ionize to yield the H^+ ion and a conjugate base. Because HCl is a strong acid, its conjugate base $\left(\mathrm{Cl}^-\right)$ is extremely weak. The chloride ion is incapable of accepting the H^+ ion and becoming HCl again. In general, the stronger the acid, the weaker its conjugate base. Likewise, the weaker the acid, the stronger its conjugate base.

Table 11.12.1: Relative Strengths of Acids and their Conjugate Bases

Acid	Conjugate Base
Strong Acids	
HCl (hydrochloric acid) (strongest)	Cl^- (chloride ion) (weakest)
$\mathrm{H_2SO_4}$ (sulfuric acid)	$\mathrm{HSO_4^-}$ (hydrogen sulfate ion)
$\mathrm{HNO_3}$ (nitric acid)	$\mathrm{NO_3^-}$ (nitrate ion)
Weak Acids	
$\mathrm{H_3PO_4}$ (phosphoric acid)	$\mathrm{H_2PO_4^-}$ (dihydrogen phosphate ion)
$\mathrm{CH_3COOH}$ (acetic acid)	$\mathrm{CH_3COO^-}$ (acetate ion)
$\mathrm{H_2CO_3}$ (carbonic acid)	$\mathrm{HCO_3^-}$ (hydrogen carbonate ion)
HCN (hydrocyanic acid) (weakest)	CN^- (cyanide ion) (strongest)

Strong acids are 100% ionized in solution. Weak acids are only slightly ionized. Phosphoric acid is stronger than acetic acid, and so is ionized to a greater extent. Acetic acid is stronger than carbonic acid, and so on.

The Acid Ionization Constant, K_a

The ionization for a general weak acid, HA, can be written as follows:

$$\mathrm{HA}\,(aq) \rightarrow \mathrm{H}^+\,(aq) + \mathrm{A}^-\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_\mathrm{a} = \frac{\left[\mathrm{H}^+\right]\left[\mathrm{A}^-\right]}{[\mathrm{HA}]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution. Therefore, the numerical value of K_a is a reflection of the strength of the acid. Weak acids with relatively higher K_a values are stronger than acids with relatively lower K_a values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the K_a value approaches infinity. For this reason, K_a values are generally reported for weak acids only.

The table below is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table 11.12.2: Acid Ionization Constants at 25^oC

Name of Acid	**Ionization Equation**	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$ $HSO_4 \rightleftharpoons H^+ + SO_4^{2-}$	very large 1.3×10^{-2}
Oxalic acid	$H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$ $HC_2O_4 \rightleftharpoons H^+ + C_2O_4^{2-}$	6.5×10^{-2} 6.1×10^{-5}
Phosphoric acid	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ $H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2--}$ $HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$	7.5×10^{-3} 6.2×10^{-8} 4.8×10^{-13}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.2×10^{-7} 4.8×10^{-11}
Hydrocyanic acid	$HCN \rightleftharpoons H^+ + CN^-$	4.9×10^{-10}

Summary

- A strong acid is an acid which is completely ionized in an aqueous solution.
- A weak acid is an acid that ionizes only slightly in an aqueous solution.
- The acid ionization constant (K_a) is defined.

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11.13: Strong and Weak Bases and Base Ionization Constant

All the complex electronics and apparatuses in a space shuttle generate heat, as do the astronauts. The shuttles have a complex arrangement of systems to dissipate that heat into outer space. One of the components of this system is a series of coils filled with ammonia that are located on the outside of the shuttle. Ammonia absorbs the heat and then releases it into space as the gas circulates through the coils. This approach is both inexpensive and effective.

Strong and Weak Bases and Base Ionization Constant, K_b

As with acids, bases can either be strong or weak, depending on the extent of their ionization. A **strong base** is a base that ionizes completely in an aqueous solution. The most common strong bases are soluble metal hydroxide compounds such as potassium hydroxide. Some metal hydroxides are not as strong, simply because they are not as soluble. Calcium hydroxide is only slightly soluble in water, but the portion that does dissolve also dissociates into ions.

A **weak base** is a base that ionizes only slightly in an aqueous solution. Recall that a base can be defined as a substance that accepts a hydrogen ion from another substance. When a weak base such as ammonia is dissolved in water, it accepts an H^+ ion from water, forming the hydroxide ion and the conjugate acid of the base, the ammonium ion.

$$\text{NH}_3\left(aq\right)+\text{H}_2\text{O}\left(l\right)\rightleftharpoons\text{NH}_4^+\left(aq\right)+\text{OH}^-\left(aq\right)$$

The equilibrium greatly favors the reactants and the extent of ionization of the ammonia molecule is very small.

An equilibrium expression can be written for the reactions of weak bases with water. Because the concentration of water is extremely large and virtually constant, the water is not included in the expression. A **base ionization constant** $\left(K_\text{b}\right)$ is the equilibrium constant for the ionization of a base. For ammonia, the expression is:

$$K_\text{b}=\frac{\left[\text{NH}_4^+\right]\left[\text{OH}^-\right]}{\left[\text{NH}_3\right]}$$

The numerical value of K_b is a reflection of the strength of the base. Weak bases with relatively high K_b values are stronger than bases with relatively low K_b values. The table below is a listing of base ionization constants for several weak bases.

Table 11.13.1: Base Ionization Constants at 25^oC

Name of Base	Ionization Equation	K_b
Methylamine	$\text{CH}_3\text{NH}_2+\text{H}_2\text{O}\rightleftharpoons\text{CH}_3\text{NH}_3^++\text{OH}^-$	5.6×10^{-4}
Ammonia	$\text{NH}_3+\text{H}_2\text{O}\rightleftharpoons\text{NH}_4^++\text{OH}^-$	1.8×10^{-5}
Pyridine	$\text{C}_5\text{H}_5\text{N}+\text{H}_2\text{O}\rightleftharpoons\text{C}_5\text{H}_5\text{NH}^++\text{OH}^-$	1.7×10^{-9}
Acetate ion	$\text{CH}_3\text{COO}^-+\text{H}_2\text{O}\rightleftharpoons\text{CH}_3\text{COOH}+\text{OH}^-$	5.6×10^{-10}
Fluoride ion	$\text{F}^-+\text{H}_2\text{O}\rightleftharpoons\text{HF}+\text{OH}^-$	1.4×10^{-11}
Urea	$\text{H}_2\text{NCONH}_2+\text{H}_2\text{O}\rightleftharpoons\text{H}_2\text{NCONH}_3^++\text{OH}^-$	1.5×10^{-14}

Notice that the conjugate base of a weak acid is also a strong base. For example, the acetate ion has a small tendency to accept a hydrogen ion from water to form acetic acid and the hydroxide ion.

Summary

- A strong base is a base that ionizes completely in an aqueous solution.
- A weak base is a base that ionizes only slightly in an aqueous solution.
- A base ionization constant $\left(K_\text{b}\right)$ is the equilibrium constant for the ionization of a base.
- The conjugate base of a weak acid is also a strong base.

11.14: Calculating Acid and Base Dissociation Constants

The pH meter was invented because Florida orange growers needed a way to test the acidity of their fruit. The first meter was invented by Arnold Beckman, who went on to form Beckman Instruments. Beckman's business was very successful, and he used much of his fortune to fund science education and research. The Beckman family donated $40 million to build the Beckman Institute at the University of Illinois.

Calculating K_a and K_b

The numerical value of K_a and K_b can be determined from an experiment. A solution of known concentration is prepared and its pH is measured with an instrument called a **pH meter**.

Figure 11.14.1: A pH meter is a laboratory device that provides quick, accurate measurements of the pH of solutions. (CC BY-NC; CK-12)

Example 11.14.1

A $0.500\ \mathrm{M}$ solution of formic acid is prepared and its pH is measured to be 2.04. Determine the K_a for formic acid.

Solution

Step 1: List the known values and plan the problem.

Known

- Initial $[\mathrm{HCOOH}] = 0.500\ \mathrm{M}$
- $\mathrm{pH} = 2.04$

Unknown

First, the pH is used to calculate the $[\mathrm{H}^+]$ at equilibrium. An ICE table is set up in order to determine the concentrations of HCOOH and HCOO^- at equilibrium. All concentrations are then substituted into the K_a expression and the K_a value is calculated.

Step 2: Solve.

$$[\mathrm{H}^+] = 10^{-\mathrm{pH}} = 10^{-2.04} = 9.12 \times 10^{-3}\ \mathrm{M}$$

Since each formic acid molecule that ionizes yields one H^+ ion and one formate ion (HCOO^-), the concentrations of H^+ and HCOO^- are equal at equilibrium. We assume that the initial concentrations of each ion are zero, resulting in the following ICE table.

	$HCOOH$	H^+	$HCOO^-$
Initial	0.500	0	0
Change	-9.12×10^{-3}	$+9.12 \times 10^{-3}$	$+9.12 \times 10^{-3}$
Equilibrium	0.491	9.12×10^{-3}	9.12×10^{-3}

Now, substituting into the K_a expression gives:

$$K_a = \frac{\left[H^+\right]\left[HCOO^-\right]}{[HCOOH]} = \frac{\left(9.12 \times 10^{-3}\right)\left(9.12 \times 10^{-3}\right)}{0.491} = 1.7 \times 10^{-4}$$

Step 3: Think about your result.

The value of K_a is consistent with that of a weak acid. Two significant figures are appropriate for the answer, since there are two digits after the decimal point in the reported pH.

Similar steps can be taken to determine the K_b of a base. For example, a 0.750 M solution of the weak base ethylamine $\left(C_2H_5NH_2\right)$ has a pH of 12.31.

$$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$$

Since one of the products of the ionization reaction is the hydroxide ion, we need to first find the $\left[OH^-\right]$ at equilibrium. The pOH is $14 - 12.31 = 1.69$. The $\left[OH^-\right]$ is then found from $10^{-1.69} = 2.04 \times 10^{-2} \text{ M}$. The ICE table is then set up as shown below.

	$C_2H_5NH_2$	$C_2H_5NH_3^+$	OH^-
Initial	0.750	0	0
Change	-2.04×10^{-2}	$+2.04 \times 10^{-2}$	$+2.04 \times 10^{-2}$
Equilibrium	0.730	2.04×10^{-2}	2.04×10^{-2}

Substituting into the K_b expression yields the K_b for ethylamine.

$$K_b = \frac{\left[C_2H_5NH_3^+\right]\left[OH^-\right]}{\left[C_2H_5NH_2\right]} = \frac{\left(2.04 \times 10^{-2}\right)\left(2.04 \times 10^{-2}\right)}{0.730} = 5.7 \times 10^{-4}$$

Summary

11.15: Calculating pH of Weak Acid and Base Solutions

Bees are beautiful creatures that help plants flourish. They carry pollen from one plant to another to facilitate plant growth and development. However, they can be troublesome when they sting! For those who are allergic to bee venom, this can be a serious, life-threatening problem. For all other humans, it can be a painful experience. When stung by a bee, one first-aid treatment is to apply a paste of baking soda (sodium bicarbonate) to the stung area. This weak base helps with the itching and swelling that accompanies the bee sting.

Calculating pH of Weak Acid and Base Solutions

The K_a and K_b values have been determined for a great many acids and bases, as shown in Tables 21.12.2 and 21.13.1. These can be used to calculate the pH of any solution of a weak acid or base whose ionization constant is known.

Example 11.15.1

Calculate the pH of a $2.00\ \mathrm{M}$ solution of nitrous acid $(\mathrm{HNO_2})$. The K_a for nitrous acid is 4.5×10^{-4}.

Solution

Step 1: List the known values and plan the problem.

Known

- Initial $[\mathrm{HNO_2}] = 2.00\ \mathrm{M}$
- $K_a = 4.5 \times 10^{-4}$

Unknown

First, an ICE table is set up with the variable x used to signify the change in concentration of the substance due to ionization of the acid. Then the K_a expression is used to solve for x and calculate the pH.

Step 2: Solve.

	$\mathrm{HNO_2}$	$\mathrm{H^+}$	$\mathrm{NO_2^-}$
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00-x$	x	x

The K_a expression and value are used to set up an equation to solve for x.

$$K_a = 4.5 \times 10^{-4} = \frac{(x)(x)}{2.00-x} = \frac{x^2}{2.00-x}$$

The quadratic equation is required to solve this equation for x. However, a simplification can be made of the fact that the extent of ionization of weak acids is small. The value of x will be significantly less than 2.00, so the "$-x$" in the denominator can be dropped.

$$\begin{aligned} 4.5 \times 10^{-4} &= \frac{x^2}{2.00-x} \approx \frac{x^2}{2.00} \\ x &= \sqrt{4.5 \times 10^{-4}(2.00)} = 2.9 \times 10^{-2}\ \mathrm{M} = [\mathrm{H^+}] \end{aligned}$$

Since the variable x represents the hydrogen-ion concentration, the pH of the solution can now be calculated.

$$\mathrm{pH}=-\log\left[\mathrm{H}^{+}\right]=-\log\left[2.9\times10^{-2}\right]=1.54$$

Step 3: Think about your result.

The pH of a $2.00\ \mathrm{M}$ solution of a strong acid would be equal to $-\log(2.00)=-0.30$. The higher pH of the $2.00\ \mathrm{M}$ nitrous acid is consistent with it being a weak acid and therefore not as acidic as a strong acid would be.

The procedure for calculating the pH of a solution of a weak base is similar to that of the weak acid in the example. However, the variable x will represent the concentration of the hydroxide ion. The pH is found by taking the negative logarithm to get the pOH, followed by subtracting from 14 to get the pH.

Summary

- The procedure for calculating the pH of a weak acid or base is illustrated.

11.16: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

Pouring concrete and working it are messy jobs. In the process, a lot of wastewater with an alkaline pH is generated. Often, regulations require that this wastewater be cleaned up at the site. One practical way to neutralize the basic pH is to bubble CO_2 into the water. The carbon dioxide forms a weak acid (carbonic acid, H_2CO_3) in solution which serves to bring the alkaline pH down to something closer to neutral.

Neutralization Reactions and Net Ionic Equations for Neutralization Reactions

A **neutralization reaction** is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water. The aqueous sodium chloride that is produced in the reaction is called a salt. A **salt** is an ionic compound composed of a cation from a base and an anion from an acid. A salt is essentially any ionic compound that is neither an acid nor a base.

Strong Acid-Strong Base Reactions

When equal amounts of a strong acid such as hydrochloric acid are mixed with a strong base such as sodium hydroxide, the result is a neutral solution. The products of the reaction do not have the characteristics of either an acid or a base. The balanced molecular equation is:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Chemical reactions occurring in aqueous solution are more accurately represented with a net ionic equation. The full ionic equation for the neutralization of hydrochloric acid by sodium hydroxide is written as follows:

$$\mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{H_2O}\,(l)$$

Since the acid and base are both strong, they are fully ionized and so are written as ions, as is the $NaCl$ formed as a product. The sodium and chloride ions are spectator ions in the reaction, leaving the following as the net ionic reaction.

$$\mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l)$$

All neutralization reactions of a strong acid with a strong base simplify to the net ionic reaction of hydrogen ion combining with hydroxide ion to produce water.

What if the acid is a diprotic acid such as sulfuric acid? The balanced molecular equation now involves a 1:2 ratio between acid and base.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + \mathrm{H_2O}\,(l)$$

In order for the reaction to be a full neutralization, twice as many moles of $NaOH$ must react with the H_2SO_4. The sodium sulfate salt is soluble, and so the net ionic reaction is again the same. Different mole ratios occur for other polyprotic acids or bases with multiple hydroxides such as $Ca(OH)_2$.

Reactions Involving a Weak Acid or Weak Base

Reactions where at least one of the components is weak do not generally result in a neutral solution. The reaction between weak nitrous acid and strong potassium hydroxide is shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{KOH}\,(aq) \rightarrow \mathrm{KNO_2}\,(aq) + \mathrm{H_2O}\,(l)$$

In order to write the net ionic equation, the weak acid must be written as a molecule since it does not ionize to a great extent in water. The base and the salt are fully dissociated.

$$\mathrm{HNO_2}\,(aq) + \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{NO_2^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The only spectator ion is the potassium ion, resulting in the net ionic equation:

$$\mathrm{HNO_2}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{NO_2^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The strong hydroxide ion essentially "forces" the weak nitrous acid to become ionized. The hydrogen ion from the acid combines with the hydroxide ion to form water, leaving the nitrite ion as the other product. The resulting solution is not neutral ($\mathrm{pH} = 7$), but instead is slightly basic.

Reactions can also involve a weak base and strong acid, resulting in a solution that is slightly acidic. The molecular and net ionic equations for the reaction of hydrochloric acid and ammonia are shown below.

$$\begin{array}{ll} \mathrm{HCl}\,(aq)+\mathrm{NH_3}\,(aq)\rightarrow\mathrm{NH_4Cl}\,(aq) & \\ \mathrm{H^+}\,(aq)+\mathrm{NH_3}\,(aq)\rightarrow\mathrm{NH_4^+}\,(aq) & \left(\mathrm{Cl^-}\text{ is a spectator ion}\right) \end{array}$$

Reactions between acids and bases that are both weak may result in solutions that are neutral, acidic, or basic.

Summary

- A neutralization reaction is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water.
- Equations for acid-base neutralizations are given.
- Net ionic equations for neutralization reactions are given.

11.17: Titration Experiment

A large amount of current research involves the development of biodiesel fuels. Often this material can be made from used vegetable oils. The vegetable oil is treated with lye to create the biofuel. In the oils is a variable amount of acid that needs to be determined, so that the workers will know how much lye to add to make the final fuel. Before the lye is added, the native vegetable oil is titrated to find out how much free acid is present. Then, the amount of lye added can be adjusted to take into account the amount needed to neutralize these free acids.

Titration Experiment

In the neutralization of hydrochloric acid by sodium hydroxide, the mole ratio of acid to base is 1:1.

$$\mathrm{HCl}\,(aq)+\mathrm{NaOH}\,(aq)\rightarrow\mathrm{NaCl}\,(aq)+\mathrm{H_2O}\,(l)$$

One mole of HCl would be fully neutralized by one mole of NaOH. If instead the hydrochloric acid was reacted with barium hydroxide, the mole ratio would be 2:1.

$$2\mathrm{HCl}\,(aq)+\mathrm{Ba(OH)_2}\,(aq)\rightarrow\mathrm{BaCl_2}\,(aq)+2\mathrm{H_2O}\,(l)$$

Now two moles of HCl would be required to neutralize one mole of $\mathrm{Ba(OH)_2}$. The mole ratio ensures that the number of moles of $\mathrm{H^+}$ ions supplied by the acid is equal to the number of $\mathrm{OH^-}$ ions supplied by the base. This must be the case for neutralization to occur. The **equivalence point** is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

In the laboratory, it is useful to have an experiment where the unknown concentration of an acid or a base can be determined. This can be accomplished by performing a controlled neutralization reaction. A **titration** is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. Many titrations are acid-base neutralization reactions, though other types of titrations can also be performed.

In order to perform an acid-base titration, the chemist must have a way to visually detect that the neutralization reaction has occurred. An **indicator** is a substance that has a distinctly different color when in an acidic or basic solution. A commonly used indicator for strong acid-strong base titrations is phenolphthalein. Solutions in which a few drops of phenolphthalein have been added turn from colorless to brilliant pink as the solution turns from acidic to basic. The steps in a titration reaction are outlined below.

1. A measured volume of an acid of unknown concentration is added to an Erlenmeyer flask.
2. Several drops of an indicator are added to the acid and mixed by swirling the flask.
3. A buret is filled with a base solution of known molarity.
4. The stopcock of the buret is opened and base is slowly added to the acid, while the flask is constantly swirled to ensure mixing. The stopcock is closed at the exact point at which the indicator just changes color.

The **standard solution** is the solution in a titration whose concentration is known. In the titration described above, the base solution is the standard solution. It is very important in a titration to add the solution from the buret slowly, so that the point at which the indicator changes color can be found accurately. The **end point** of a titration is the point at which the indicator changes color. When phenolphthalein is the indicator, the end point will be signified by a faint pink color.

Figure 11.17.1: Phenolphthalein in basic solution. (CC BY-NC; CK-12)

Summary

- The equivalence point is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

- A titration is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. (Many titrations are acid-base neutralization reactions.)
- An indicator is a substance that has a distinctly different color when in an acidic or basic solution.
- The process for carrying out a titration is described.

11.18: Titration Calculations

The manufacture of soap requires a number of chemistry techniques. One necessary piece of information is the saponification number. This is the amount of base needed to hydrolyze a certain amount of fat to produce the free fatty acids that are an essential part of the final product. The fat is heated with a known amount of base (usually NaOH or KOH). After hydrolysis is complete, the leftover base is titrated to determine how much was needed to hydrolyze the fat sample.

Titration Calculations

At the equivalence point in a neutralization, the moles of acid are equal to the moles of base.

$$\text{moles acid} = \text{moles base}$$

Recall that the molarity (M) of a solution is defined as the moles of the solute divided by the liters of solution (L). So the moles of solute are therefore equal to the molarity of a solution multiplied by the volume in liters.

$$\text{moles solute} = \mathrm{M} \times \mathrm{L}$$

We can then set the moles of acid equal to the moles of base.

$$\mathrm{M}_A \times \mathrm{V}_A = \mathrm{M}_B \times \mathrm{V}_B$$

M_A is the molarity of the acid, while M_B is the molarity of the base. V_A and V_B are the volumes of the acid and base, respectively.

Suppose that a titration is performed and $20.70\ \mathrm{mL}$ of $0.500\ \mathrm{M}\ \mathrm{NaOH}$ is required to reach the end point when titrated against $15.00\ \mathrm{mL}$of HCl of unknown concentration. The above equation can be used to solve for the molarity of the acid.

$$\mathrm{M}_A = \frac{\mathrm{M}_B \times \mathrm{V}_B}{\mathrm{V}_A} = \frac{0.500\ \mathrm{M} \times 20.70\ \mathrm{mL}}{15.00\ \mathrm{mL}} = 0.690\ \mathrm{M}$$

The higher molarity of the acid compared to the base in this case means that a smaller volume of the acid is required to reach the equivalence point.

The above equation works only for neutralizations in which there is a 1:1 ratio between the acid and the base. The example below demonstrates the technique to solve a titration problem for a titration of sulfuric acid with sodium hydroxide.

Example 11.18.1

In a titration of sulfuric acid against sodium hydroxide, $32.20\ \mathrm{mL}$ of $0.250\ \mathrm{M}\ \mathrm{NaOH}$ is required to neutralize $26.60\ \mathrm{mL}$ of $\mathrm{H_2SO_4}$. Calculate the molarity of the sulfuric acid.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + 2\mathrm{H_2O}\,(l)$$

Solution

Step 1: List the known values and plan the problem.

- Molarity $\mathrm{NaOH} = 0.250\ \mathrm{M}$
- Volume $\mathrm{NaOH} = 32.20\ \mathrm{mL}$
- Volume $\mathrm{H_2SO_4} = 26.60\ \mathrm{mL}$

Unknown

First determine the moles of NaOH in the reaction. From the mole ratio, calculate the moles of $\mathrm{H_2SO_4}$ that reacted. Finally, divide the moles of $\mathrm{H_2SO_4}$ by its volume to get the molarity.

Step 2: Solve.

$$\text{mol NaOH} = \text{M} \times \text{L} = 0.250 \text{ M} \times 0.03220 \text{ L} = 8.05 \times 10^{-3} \text{ mol NaOH}$$

$$8.05 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 4.03 \times 10^{-3} \text{ mol H}_2\text{SO}_4$$

$$\frac{4.03 \times 10^{-3} \text{ mol H}_2\text{SO}_4}{0.02660 \text{ L}} = 0.151 \text{ M H}_2\text{SO}_4$$

Step 3: Think about your result.

The volume of H_2SO_4 required is smaller than the volume of $NaOH$ because of the two hydrogen ions contributed by each molecule.

Summary

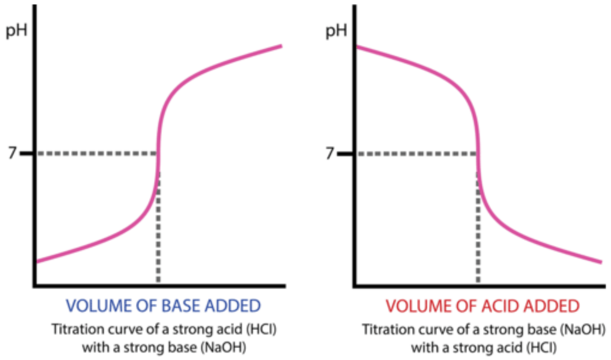
- The process of calculating concentration from titration data is described and illustrated.

11.19: Titration Curves

The x-y plot that we know of as a graph was the brainchild of the French mathematician-philosopher Rene Descartes (1596-1650). His studies in mathematics led him to develop what was known as "Cartesian geometry", including the concept of today's graphs. The coordinates are often referred to as Cartesian coordinates.

Titration Curves

As base is added to acid at the beginning of a titration, the pH rises very slowly. Nearer to the equivalence point, the pH begins to rapidly increase. If the titration is a strong acid with a strong base, the pH at the equivalence point is equal to 7. A bit past the equivalence point, the rate of change of the pH again slows down. A **titration curve** is a graphical representation of the pH of a solution during a titration. The figure below shows two different examples of a strong acid-strong base titration curve. On the left is a titration in which the base is added to the acid, and so the pH progresses from low to high. On the right is a titration in which the acid is added to the base. In this case, the pH starts out high and decreases during the titration. In both cases, the equivalence point is reached when the moles of acid and base are equal and the pH is 7. This also corresponds to the color change of the indicator.

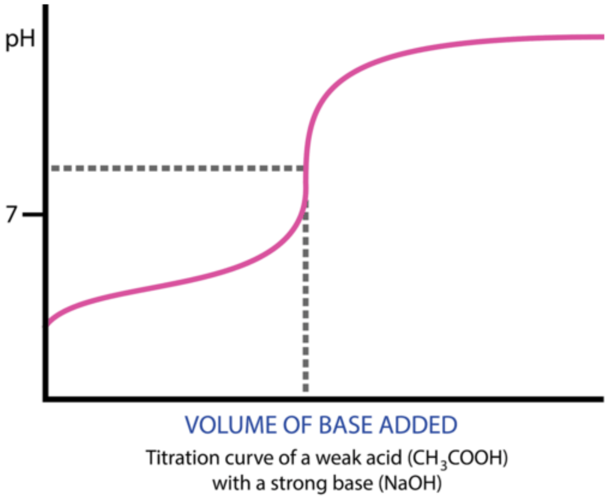


Titration curve of a strong acid (HCl) with a strong base (NaOH)

Titration curve of a strong base (NaOH) with a strong acid (HCl)

Figure 11.19.1: A titration curve shows the pH changes that occur during the titration of an acid with a base. On the left, base is being added to acid. On the right, acid is being added to base. In both cases, the equivalence point is at pH 7. (CC BY-NC; CK-12)

Titration curves can also be generated in the case of a weak acid-strong base titration or a strong acid-weak base titration. The general shape of the titration curve is the same, but the pH at the equivalence point is different. In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point. In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.



Titration curve of a weak acid (CH_3COOH) with a strong base (NaOH)

Figure 11.19.2: Titration curve of weak acid and strong base. (CC BY-NC; CK-12)

Summary

- A titration curve is a graphical representation of the pH of a solution during a titration.
- In a strong acid-strong base titration, the equivalence point is reached when the moles of acid and base are equal and the pH is 7.
- In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point.

- In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.

11.20: Indicators

"Boil Them Cabbage Down" is an old bluegrass song. Many people enjoy the music, but chemistry students also enjoy the product of boiled cabbage. Extracting the anthocyanin dye from cabbage leaves with boiling water gives a solution that is red when acidic, purple when neutral, and green to yellow when basic.

Indicators

An **acid-base indicator** is a substance that displays different colors when in the presence of an acid or a base. How does that work? An indicator is a weak acid that ionizes within a known pH range, usually about 2 pH units. We can represent the protonated form of the indicator molecule as HIn and the deprotonated form as In^-. The following equilibrium exists for the indicator:

$$\mathrm{HIn}\,(aq) \rightleftharpoons \mathrm{H}^+\,(aq) + \mathrm{In}^-\,(aq)$$

According to Le Chatelier's principle, the addition of H^+ ions (as in a low pH solution) drives the equilibrium to the left and the protonated HIn predominates. The addition of OH^- (as in a high pH solution) decreases the H^+ concentration and drives the equilibrium to the right and the deprotonated In^- predominates. To be useful as an indicator, the two forms must be different colors. In the case of phenolphthalein, the protonated form is colorless, while the deprotonated form is pink. The figure below shows a variety of acid-base indicators that can be used in titration experiments.

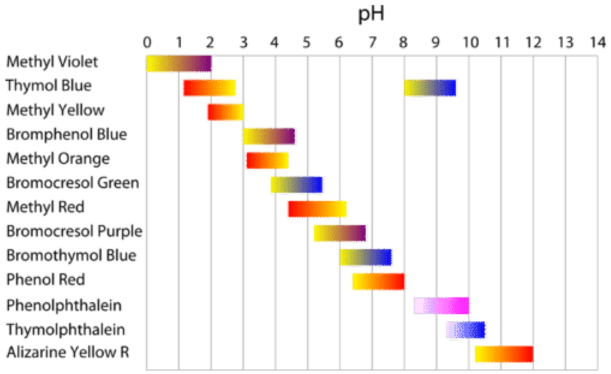


Figure 11.20.1: pH indicators. (CC BY-NC; CK-12)

Depending on the pH at the equivalence point, the appropriate indicator must be chosen. For example, bromphenol blue has a yellow color below a pH of about 3 and a blue-violet color above a pH of about 4. Bromphenol blue would not be a good choice as the indicator for a strong acid-strong base titration, because the pH is 7 at the equivalence point. Instead, it could be used for a strong acid-weak base titration, where the pH at the equivalence point is lower.

Most indicators have two colored forms. Universal indicator displays the entire rainbow of colors from low pH to high pH (see figure below). Universal indicator is used to make pH paper, which can be used to quickly test solutions for their approximate pH.

Figure 11.20.2: Universal indicator tape. (CC BY-NC; CK-12)

Summary

- An acid-base indicator is a substance that displays different colors when in the presence of an acid or a base.
- Examples of acid-base indicators are given.

11.21: Hydrolysis of Salts - Equations

Baking seems easy with all the pre-mixed items available. ("just add water and stir"). However, there is a good amount of chemistry involved in baking with ingredients that are measured out. One important ingredient is baking powder. The fluffiness in the final product of a non-yeast recipe is usually due to the carbon dioxide formed by baking powder. One popular brand uses a mix of sodium bicarbonate and sodium aluminum sulfate to produce the $\mathrm{CO_2}$. The reaction is:

$$\mathrm{3NaHCO_3 + NaAl(SO_4)_2 \rightarrow Al(OH)_3 + 2Na_2SO_4 + 3CO_2}$$

If all goes well, the biscuits rise, the pancakes are fluffy, and everybody is happy.

Hydrolysis of Salts: Equations

A salt is an ionic compound that is formed when an acid and a base neutralize each other. While it may seem that salt solutions are always neutral, they can frequently be either acidic or basic.

Consider the salt formed when the weak acid hydrofluoric acid is neutralized by the strong base sodium hydroxide. The molecular and net ionic equations are shown below.

$$\begin{gathered}\mathrm{HF}\,(aq)+\mathrm{NaOH}\,(aq)\rightarrow\mathrm{NaF}\,(aq)+\mathrm{H_2O}\,(l)\\ \mathrm{HF}\,(aq)+\mathrm{OH^-}\,(aq)\rightarrow\mathrm{F^-}\,(aq)+\mathrm{H_2O}\,(l)\end{gathered}$$

Since sodium fluoride is soluble, the sodium ion is a spectator ion in the neutralization reaction. The fluoride ion is capable of reacting, to a small extent, with water, accepting a proton.

$$\mathrm{F^-}\,(aq)+\mathrm{H_2O}\,(l)\rightleftharpoons\mathrm{HF}\,(aq)+\mathrm{OH^-}\,(aq)$$

The fluoride ion is acting as a weak Brønsted-Lowry base. The hydroxide ion that is produced as a result of the above reaction makes the solution slightly basic. **Salt hydrolysis** is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.

Salts That Form Basic Solutions

When solid sodium fluoride is dissolved into water, it completely dissociates into sodium ions and fluoride ions. The sodium ions do not have any capability of hydrolyzing, but the fluoride ions hydrolyze to produce a small amount of hydrofluoric acid and hydroxide ion.

$$\mathrm{F^-}\,(aq)+\mathrm{H_2O}\,(l)\rightleftharpoons\mathrm{HF}\,(aq)+\mathrm{OH^-}\,(aq)$$

Salts that are derived from the neutralization of a weak acid (HF) by a strong base (NaOH) will always produce salt solutions that are basic.

Salts That Form Acidic Solutions

Ammonium chloride $\left(\mathrm{NH_4Cl}\right)$ is a salt that is formed when the strong acid HCl is neutralized by the weak base $\mathrm{NH_3}$. Ammonium chloride is soluble in water. The chloride ion produced is incapable of hydrolyzing because it is the conjugate base of the strong acid HCl. In other words, the $\mathrm{Cl^-}$ ion cannot accept a proton from water to form HCl and $\mathrm{OH^-}$, as the fluoride ion did in the previous section. However, the ammonium ion is capable of reacting slightly with water, donating a proton and so acting as an acid.

$$\mathrm{NH_4^+}\,(aq)+\mathrm{H_2O}\,(l)\rightleftharpoons\mathrm{H_3O^+}\,(aq)+\mathrm{NH_3}\,(aq)$$

Salts That Form Neutral Solutions

A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7. An example is sodium chloride, formed from the neutralization of HCl by NaOH. A solution of NaCl in water has no acidic or basic properties, since neither ion is capable of hydrolyzing. Other salts that form neutral solutions include potassium nitrate $\left(\mathrm{KNO_3}\right)$ and lithium bromide (LiBr). The table below summarizes how to determine the acidity or basicity of a salt solution.

Table 11.21.1

Salt formed from:	Salt Solution
Strong acid + Strong base	Neutral

Salt formed from:	Salt Solution
Strong acid + Weak base	Acidic
Weak acid + Strong base	Basic

Salts formed from the reaction of a weak acid and a weak base are more difficult to analyze due to competing hydrolysis reactions between the cation and the anion. These salts are not considered in this chapter's concept.

Summary

- Salt hydrolysis is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.
- Salts that are derived from the neutralization of a weak acid by a strong base will always produce salt solutions that are basic.
- Salts that are derived from the neutralization of a strong acid by a weak base will always produce salt solutions that are acidic.
- A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7.

11.22: Calculating pH of Salt Solutions

Many enjoy a cool dip in a swimming pool on a hot day, but may not realize the work needed to keep that water safe and healthy. The ideal pH for a swimming pool is around 7.2. The pH will change as a result of many factors. Adjustment can be accomplished with different chemicals, depending on the tested pH. High pH can be lowered with liquid HCl (unsafe material) or sodium bisulfate. The bisulfate anion is a weak acid and can dissociate partially in solution. To increase pH, use sodium carbonate. The carbonate anion forms an equilibrium with protons that results in some formation of carbon dioxide.

Calculating pH of Salt Solutions

It is often helpful to be able to predict the effect a salt solution will have on the pH of a certain solution. Knowledge of the relevant acidity or basicity constants allows us to carry out the necessary calculations.

Example 11.22.1

If we dissolve NaF in water, we get the following equilibrium:

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

The pH of the resulting solution can be determined if the K_b of the fluoride ion is known. $20.0\ \mathrm{g}$ of sodium fluoride is dissolved in enough water to make $500.0\ \mathrm{mL}$ of solution. Calculate the pH of the solution. The K_b of the fluoride ion is 1.4×10^{-11}.

Solution:

Step 1: List the known values and plan the problem.

Known

- Mass $\mathrm{NaF} = 20.0\ \mathrm{g}$
- Molar mass $\mathrm{NaF} = 41.99\ \mathrm{g/mol}$
- Volume solution $= 0.5000\ \mathrm{L}$
- K_b of $\mathrm{F^-} = 1.4 \times 10^{-11}$

Unknown

The molarity of the $\mathrm{F^-}$ solution can be calculated from the mass, molar mass, and solution volume. Since NaF completely dissociates, the molarity of the NaF is equal to the molarity of the $\mathrm{F^-}$ ion. An ICE table (below) can be used to calculate the concentration of $\mathrm{OH^-}$ produced and then the pH of the solution.

Step 2: Solve.

$$20.0\ \mathrm{g}\ \cancel{\mathrm{NaF}} \times \frac{1\ \mathrm{mol}\ \cancel{\mathrm{NaF}}}{41.99\ \mathrm{g}\ \cancel{\mathrm{NaF}}} \times \frac{1\ \mathrm{mol\ F^-}}{1\ \mathrm{mol}\ \cancel{\mathrm{NaF}}} = 0.476\ \mathrm{mol\ F^-}$$

$$\frac{0.476\ \mathrm{mol\ F^-}}{0.5000\ \mathrm{L}} = 0.953\ \mathrm{M\ F^-}$$

$$\text{Hydrolysis equation:}\quad \mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

	F^-	HF	OH^-
Initial	0.953	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.953-x$	x	x

$$\begin{aligned} K_b &= 1.4 \times 10^{-11} = \frac{(x)(x)}{0.953-x} = \frac{x^2}{0.953-x} \approx \frac{x^2}{0.953} \\ x &= [OH^-] = \sqrt{1.4 \times 10^{-11}(0.953)} = 3.65 \times 10^{-6}\text{ M} \\ \text{pOH} &= -\log\left(3.65 \times 10^{-6}\right) = 5.44 \\ \text{pH} &= 14 - 5.44 = 8.56 \end{aligned}$$

Step 3: Think about your result.

The solution is slightly basic due to the hydrolysis of the fluoride ion.

Salts That Form Acidic Solutions

When the ammonium ion dissolves in water, the following equilibrium exists:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$$

The production of hydronium ions causes the resulting solution to be acidic. The pH of a solution of ammonium chloride can be found in a very similar way to the sodium fluoride solution in the previous example. However, since the ammonium chloride is acting as an acid, it is necessary to know the K_a of NH_4^+, which is 5.6×10^{-10}. We will find the pH of a 2.00 M solution of NH_4Cl. Because the NH_4Cl completely ionizes, the concentration of the ammonium ion is 2.00 M.

$$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

Again, an ICE table (below) is set up in order to solve for the concentration of the hydronium (or H^+) ion produced.

	NH_4^+	H^+	NH_3
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00-x$	x	x

Now substituting into the K_a expression gives:

$$\begin{aligned} K_a &= 5.6 \times 10^{-10} = \frac{x^2}{2.00-x} \approx \frac{x^2}{2.00} \\ x &= [H^+] = \sqrt{5.6 \times 10^{-10}(2.00)} = 3.3 \times 10^{-5}\text{ M} \\ \text{pH} &= -\log\left(3.3 \times 10^{-5}\right) = 4.48 \end{aligned}$$

A salt produced from a strong acid and a weak base yields a solution that is acidic.

Summary

11.23: Buffers

Diabetes mellitus is a disorder of glucose metabolism in which insulin production by the pancreas is impaired. Since insulin helps glucose enter the cells, a decrease of this hormone means that glucose cannot be used in its normal fashion. When this happens, the body begins to break down fats, producing a decrease in blood pH. Chemical systems in the body can balance this pH shift for a while, but excessive acid production can create serious problems if not corrected by administering insulin to restore normal glucose use.

Buffers

If only $1.0\ \text{mL}$ of $0.10\ \text{M}$ hydrochloric acid is added to $1.0\ \text{L}$ of pure water, the pH drops drastically from 7.0 to 4.0. This is a 1000-fold increase in the acidity of the solution. For many purposes, it is desirable to have a solution which is capable of resisting such large changes in pH when relatively small amounts of acid or base are added to it. Such a solution is called a buffer. A **buffer** is a solution of a weak acid or a base and its salt. Both components must be present for the system to act as a buffer to resist changes in pH. Commercial buffer solutions, which have a wide variety of pH values, can be obtained.

Some common buffer systems are listed in the table below.

Table 11.23.1: Some Common Buffers

Buffer System	Buffer Components	pH of buffer (equal molarities of both components)
Acetic acid/acetate ion	CH_3COOH/CH_3COO^-	4.74
Carbonic acid/hydrogen carbonate ion	H_2CO_3/HCO_3^-	6.38
Dihydrogen phosphate ion/hydrogen phosphate ion	$H_2PO_4^-/HPO_4^{2-}$	7.21
Ammonia/ammonium ion	NH_3/NH_4^+	9.25

One example of a buffer is a solution made of acetic acid (the weak acid) and sodium acetate (the salt). The pH of a buffer consisting of $0.50\ \text{M}\ CH_3COOH$ and $0.50\ \text{M}\ CH_3COONa$ is 4.74. If $10.0\ \text{mL}$ of $1.0\ \text{M}\ HCl$ is added to $1.0\ \text{L}$ of the buffer, the pH only decreases to 4.73. This ability to "soak up" the additional hydrogen ions from the HCl that was added is due to the reaction below.

$$CH_3COO^-\,(aq) + H^+\,(aq) \rightarrow CH_3COOH\,(aq)$$

Since both the acetate ion and the acetic acid were already present in the buffer, the only thing that changes is the ratio of one to the other. Small changes in that ratio have only very minor effects on the pH.

If $10.0\ \text{mL}$ of $1.0\ \text{M}\ NaOH$ were added to another $1.0\ \text{L}$ of the same buffer, the pH would only increase to 4.76. In this case the buffer takes up the additional hydroxide ions.

$$CH_3COOH\,(aq) + OH^-\,(aq) \rightarrow CH_3COO^-\,(aq) + H_2O\,(l)$$

Again the ratio of acetate ion to acetic acid changes slightly, this time causing a very small increase in the pH.

It is possible to add so much acid or base to a buffer that its ability to resist a significant change in pH is overwhelmed. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs. The buffer capacity is exceeded when the number of moles of H^+ or OH^- that are added to the buffer exceeds the number of moles of the buffer components.

Summary

- A buffer is a solution of a weak acid or a base and its salt.
- The buffer capacity is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs.
- Reactions showing how buffers regulate pH are described.

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Glossary

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Detailed Licensing

Overview

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