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Measurements
Chapter 1

Significant Figures vs. Precision

How many significant figures do these numbers contain ?

	12.000	0.105	0.0005020	3600
Significant Figures	?	?	?	?

- **Rules of Significant Figures:**
- **Nonzero integers - always count as significant figures.**
- **Zeros –**
  - Leading zeros - are zeros that precede all of the nonzero digits. They never count as significant figures.
  - Captive zeros - are zeros that fall between nonzero digits. They all count as significant figures.
  - Trailing zeros - are zeros at the right end of the number. They are significant only if the number contains a decimal point.
- **Exact numbers - such as tallies or conversion factors have unlimited number of significant figures.**

Forget memorizing these rules because there is a simpler method

•To determine the number of Significant Figures

• Pacific-Atlantic Rule

• Decimal Present vs. Decimal Absent

Place the number in the center of the map.

If the number has a decimal then determine if the analysis is going to originate from the pacific or the atlantic.

Decimal Present

For numbers with a decimal point present, draw a line starting from the pacific to the first non-zero number, all digits shown including the non-zero number are significant. i.e., 0.040050

Decimal Absent

For numbers with a decimal point absent, draw a line starting from the Atlantic (right) to the first non-zero number, all digits shown including the non-zero number are significant. i.e., 30500

0.040050

5 Significant Figures

30500

3 Significant Figures

Significant Figures example-

Use the Pacific-Atlantic trick to determine the number of significant figures for the following measurents.

	12.000	0.105	0.0005020	3600
Significant Figures	?	?	?	?

12.000 five significant figures

0.105 three significant figures

0.0005020 four significant figures

3600 two significant figures

What is the difference between Precision & Significant Figures

What is the difference between significant figures and precision? For each of the example shown, which between the two numbers contain the fewest significant figures and which is the least precise?

.				Fewest Significant figures	Least precise
a)	123	vs	1.2	1.2	123
b)	1.23	vs	1.200	1.23	1.23
c)	0.123	vs	0.00012	0.00012	0.123
d)	30	vs	3600	30	3600

The number with the fewest significant figures is the number which contains the lower amount of digits (significant digits). The least precise number, on the other hand, is the number with the largest uncertainty. For example 0.1 (or 1/10) has more uncertainty than 0.01 (or 1/100).

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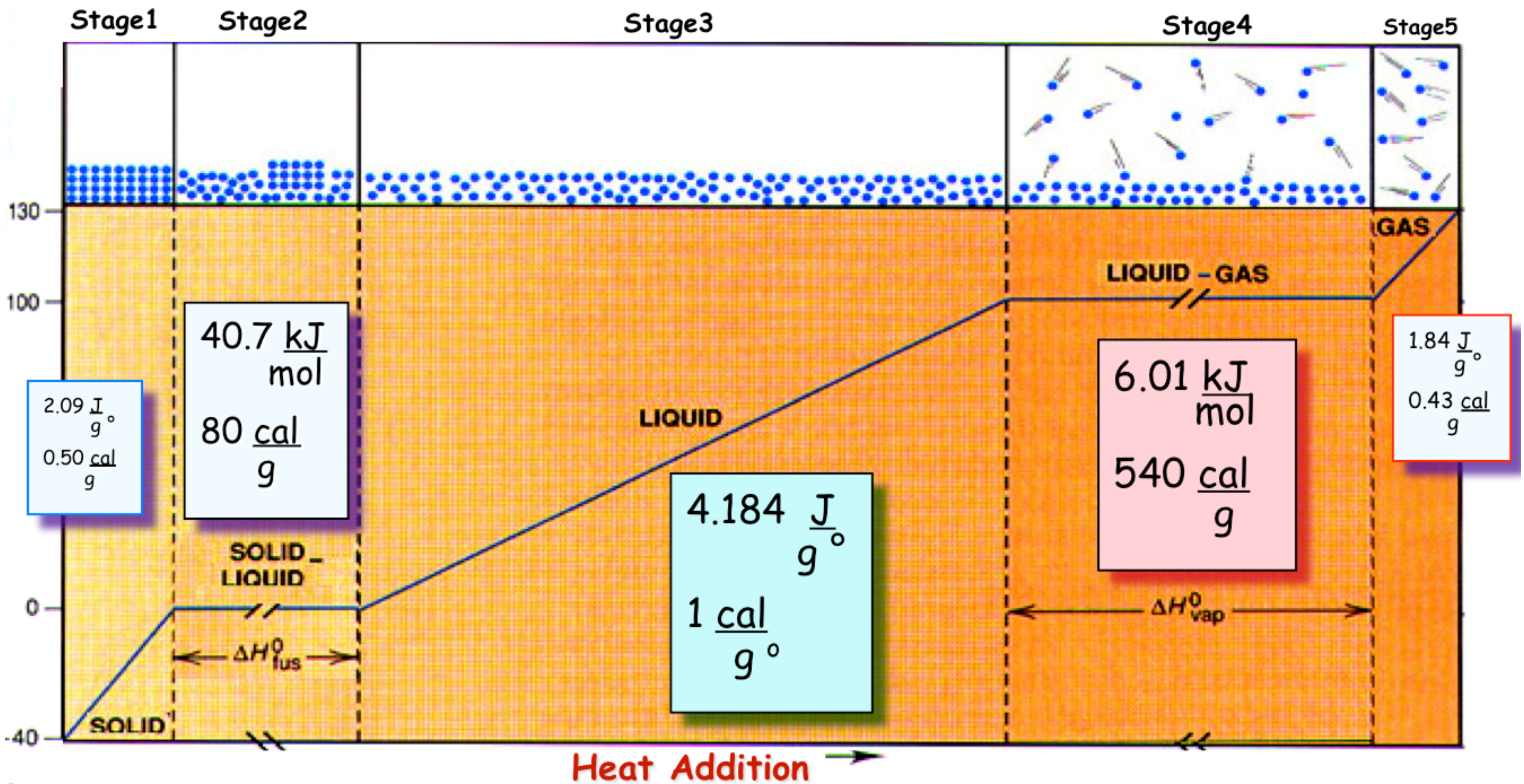
Heating-Cooling Curve
Chapter 2

Heating and Cooling Curve for H_2O

Energy Calculation for Ice to Steam and vice-versa



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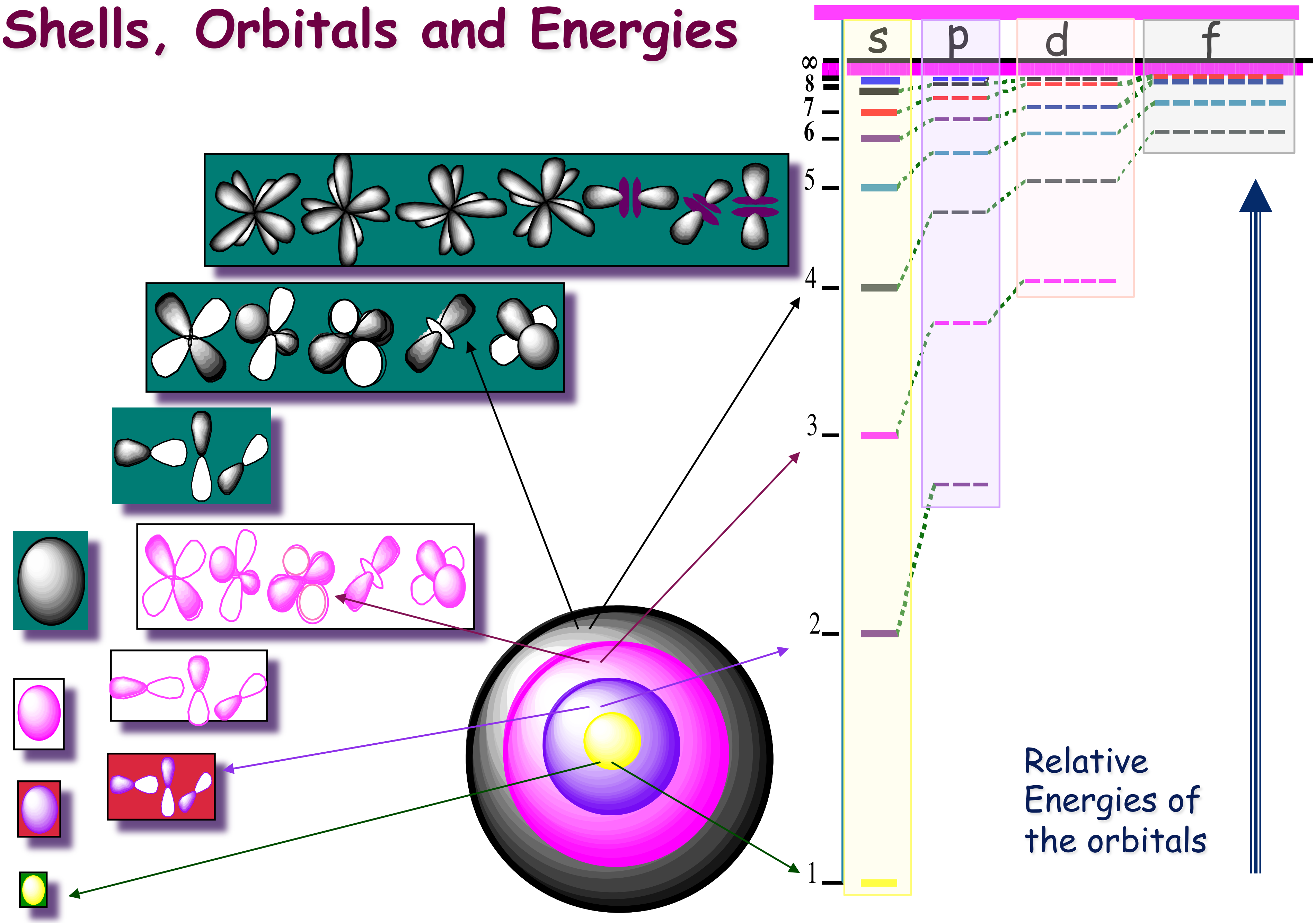
Atoms and Elements
Chapter 3

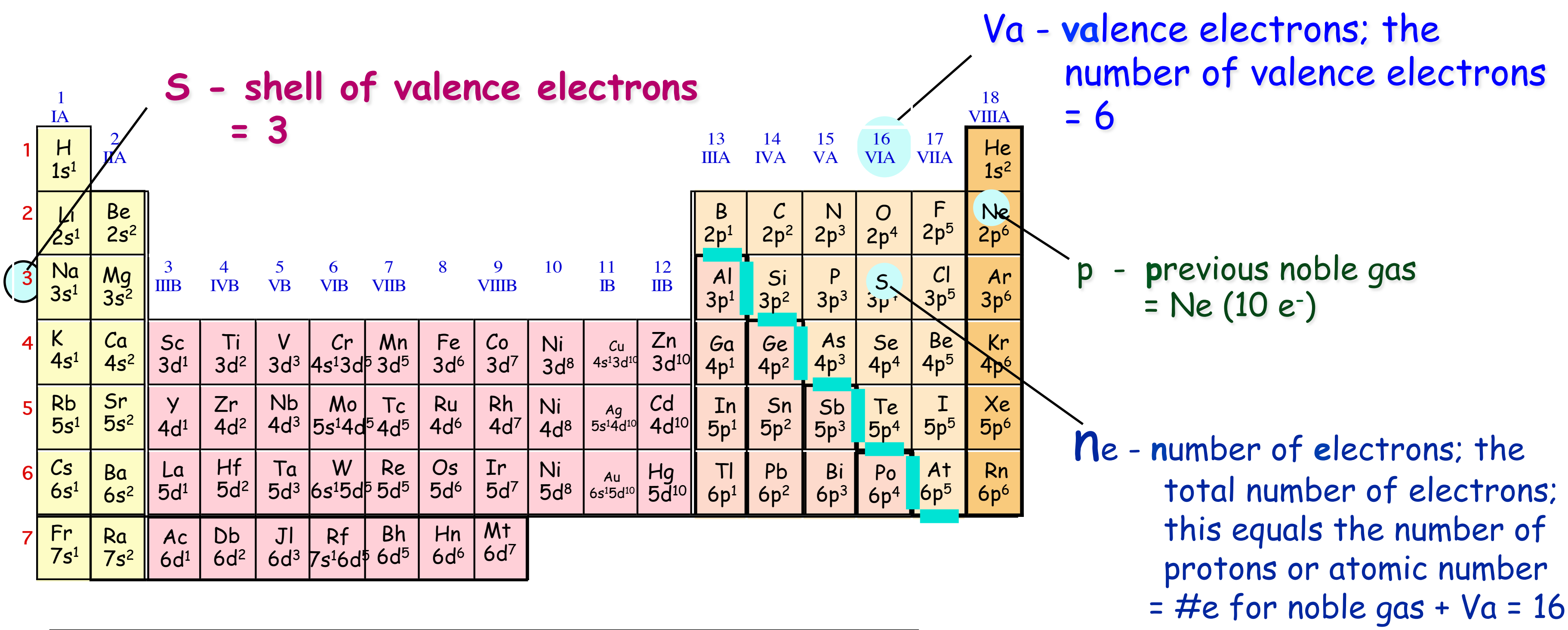
Shells, Orbitals and Energies



Supplemental Information

Noble Gas Notation for Electron Configuration.
"Shorthand Notation" (n_e-Va-s-p)

The noble gas notation for electron configuration is a way of writing the electron configuration for an element using the symbol of the noble gas to represent the "core" inner electrons. Consider the example below for sulfur below.



electron configuration for Sulfur:

S = [p] s (orbital)^{Va} = [Ne] $3s^2$ $3p^4$

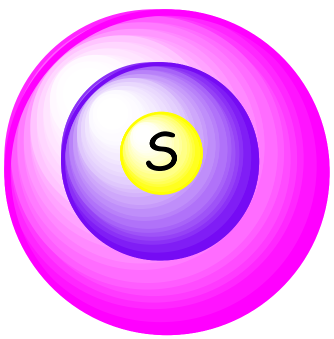
n_e = 10 + (3 + 4) = 16 Total electrons

Supplemental Information

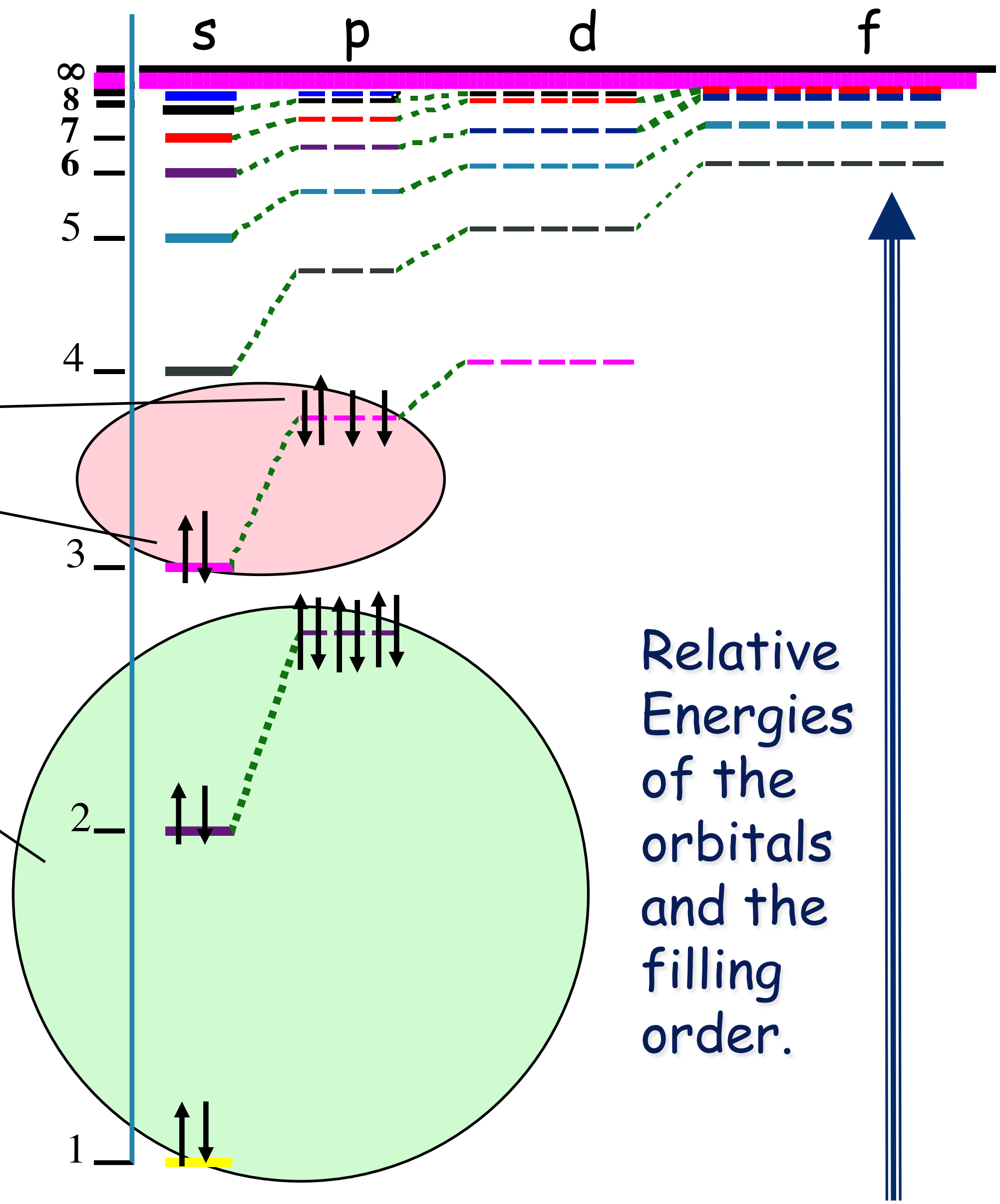
Comparing Noble gas electron configuration to energy level diagram for an atom.

•The 16 electrons for sulfur occupy the shells & orbitals of sulfur from the lowest energy to the highest.

e⁻ config for Sulfur
$S = [Ne]3s^2 3p^4$



16 total electrons



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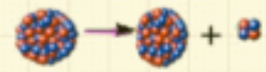
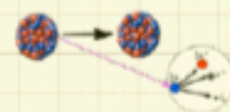
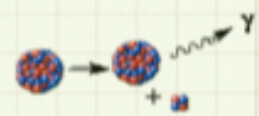
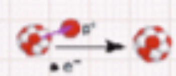
Nuclear Chemistry
Chapter 4

Nuclear Chemistry: Nomenclature

Particle	Charge	Mass (g)	Nomenclature
alpha	2+	6.64e-24	$^{4}_{2}He$ $^{4}_{2}\alpha$
beta	1-	9.11e-28	$^{0}_{-1}e$ $^{0}_{-1}\beta$
gamma	0	0	$^{0}_{0}\gamma$
proton	1+	1.673 e-24	$^{1}_{1}H$ $^{1}_{1}p$
neutron	0	1.675 e-24	$^{1}_{0}n$
electron	1-	9.11e-28	$^{0}_{-1}e$
position	1+	9.11e-28	$^{0}_{+1}e$

Mode	Emission	Decay Process	Change in A	Change in Z
α emission alpha decay	$\alpha\ (^{4}_{2}He^{2+})$	 $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}\alpha$	-4	-2
β Emission beta decay	$\beta\ (^{0}_{-1}e)$	 $^{131}_{53}I \rightarrow {}^{131}_{54}Xe + {}^{0}_{-1}\beta$	0	+1
γ Emission gamma emission	$\gamma\ (^{0}_{0}\gamma)$	 $^{215}_{84}Po \rightarrow {}^{211}_{82}Pb + {}^{4}_{2}\alpha + {}^{0}_{0}\gamma$	0	0
β⁺ Emission positron emission	$\beta^{+}\ (^{0}_{+1}e)$	 $^{11}_{6}C \rightarrow {}^{11}_{5}B + {}^{0}_{+1}\beta$	0	-1
e⁻ capture electron capture	$\beta\ (^{0}_{-1}e)$	 $^{55}_{26}I + {}^{0}_{-1}e \rightarrow {}^{55}_{25}Mn + h\nu$	0	-1

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Compounds and Their Bonds
Chapter 5

Charge of Elemental Ions

The common charge or oxidation state of many main group and some transition metals elements can be determine by the a basic knowledge of the periodic table

Common Charge for metals in Type I Ionic Compounds

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H +1																	2 He 4.0026
2	3 Li +1	4 Be 9.0122											5 B 10.811	6 C 12.0112	7 N -3	8 O -2	9 F -1	10 Ne 20.179
3	11 Na +1	12 Mg +2											13 Al +3	14 Si 28.086	15 P -3	16 S -2	17 Cl -1	18 Ar 39.948
4	19 K +1	20 Ca +2	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.54	30 Zn +2	31 Ga 65.37	32 Ge 72.59	33 As -3	34 Se -2	35 Br -1	36 Kr 83.80
5	37 Rb +1	38 Sr +2	39 Y 88.905	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc [99]	44 Ru 101.07	45 Rh 102.905	46 Pd 106.4	47 Ag +1	48 Cd +2	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te -2	53 I -1	54 Xe 131.30
6	55 Cs +1	56 Ba +2	71* Lu 174.967	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 197.0	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po [210]	85 At -1	86 Rn [222]
7	87 Fr +1	88 Ra +2	103‡ Lr [260]	104 Rf [261.11]	105 Db [262.11]	106 Sg [266.12]	107 Bh [264.12]	108 Hs [269.13]	109 Mt [268.14]	110 Ds [271]	111 Rg [272]	112 [277]		114 [289]		116 [292]		

* Lanthanide Series	57 La 138.91	58 Ce 140.115	59 Pr 140.9077	60 Nd 144.24	61 Pm (145)	62 Sm 150.368	63 Eu 151.965	64 Gd 157.25	65 Tb 158.9254	66 Dy 162.50	67 Ho 164.9303	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04
‡ Actinide Series	89 Ac [227.03]	90 Th 232.0381	91 Pa 231.0359	92 U 238.0289	93 Np 237.048	94 Pu [244]	95 Am [260]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]

Oxyions... the `ate

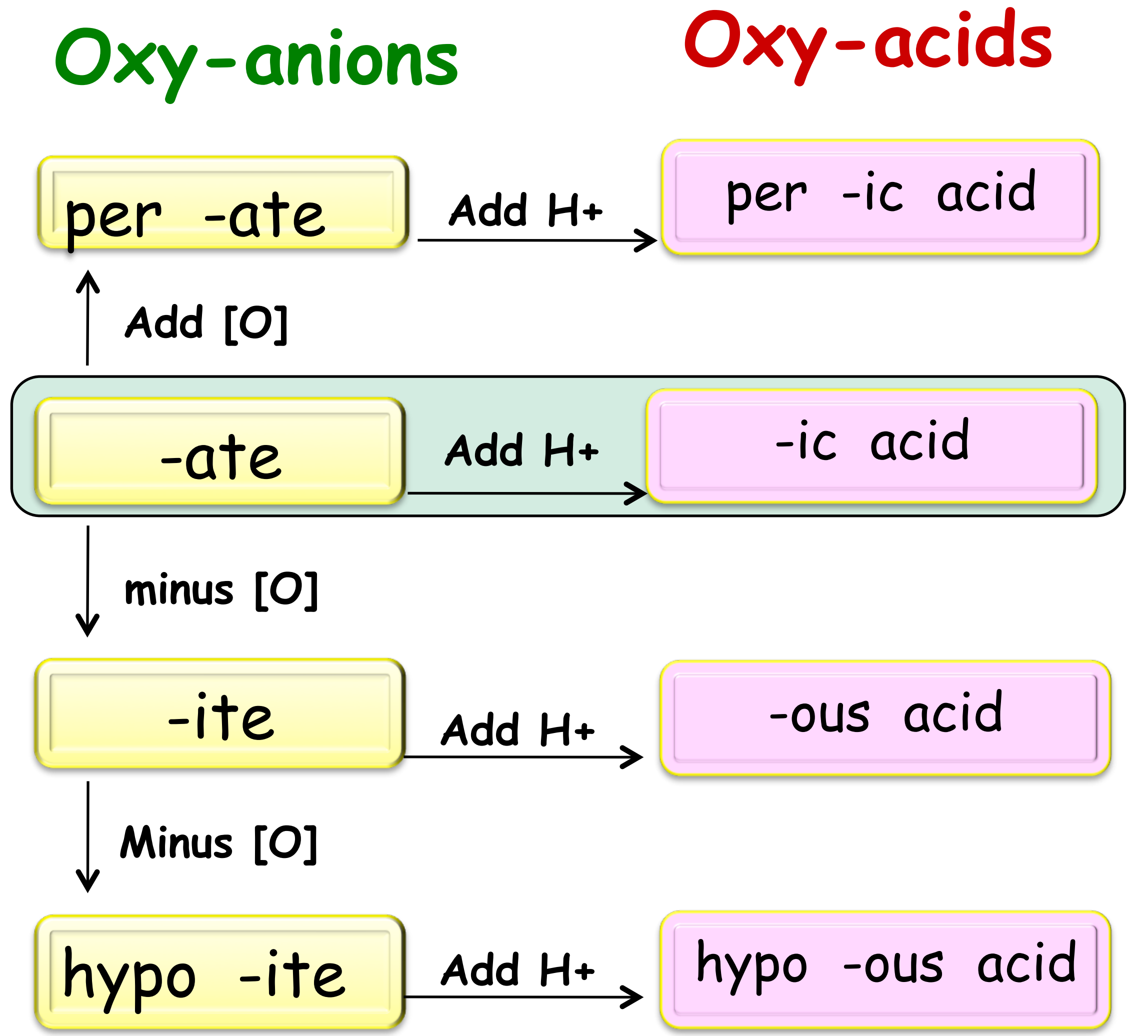
XO_3^{chg} XO_4^{chg}

BO_3^{-3}	CO_3^{-2}	NO_3^{-}	O	F	Ne
Al	Si	PO_4^{-3}	SO_4^{-2}	ClO_3^{-}	Ar
Ga	Ge	AsO_4^{-3}	SeO_4^{-2}	BrO_3^{-}	Kr
In	Sn	Sb	TeO_4^{-2}	IO_3^{-}	Xe
Tl	Pb	Bi	Po	At	Rn

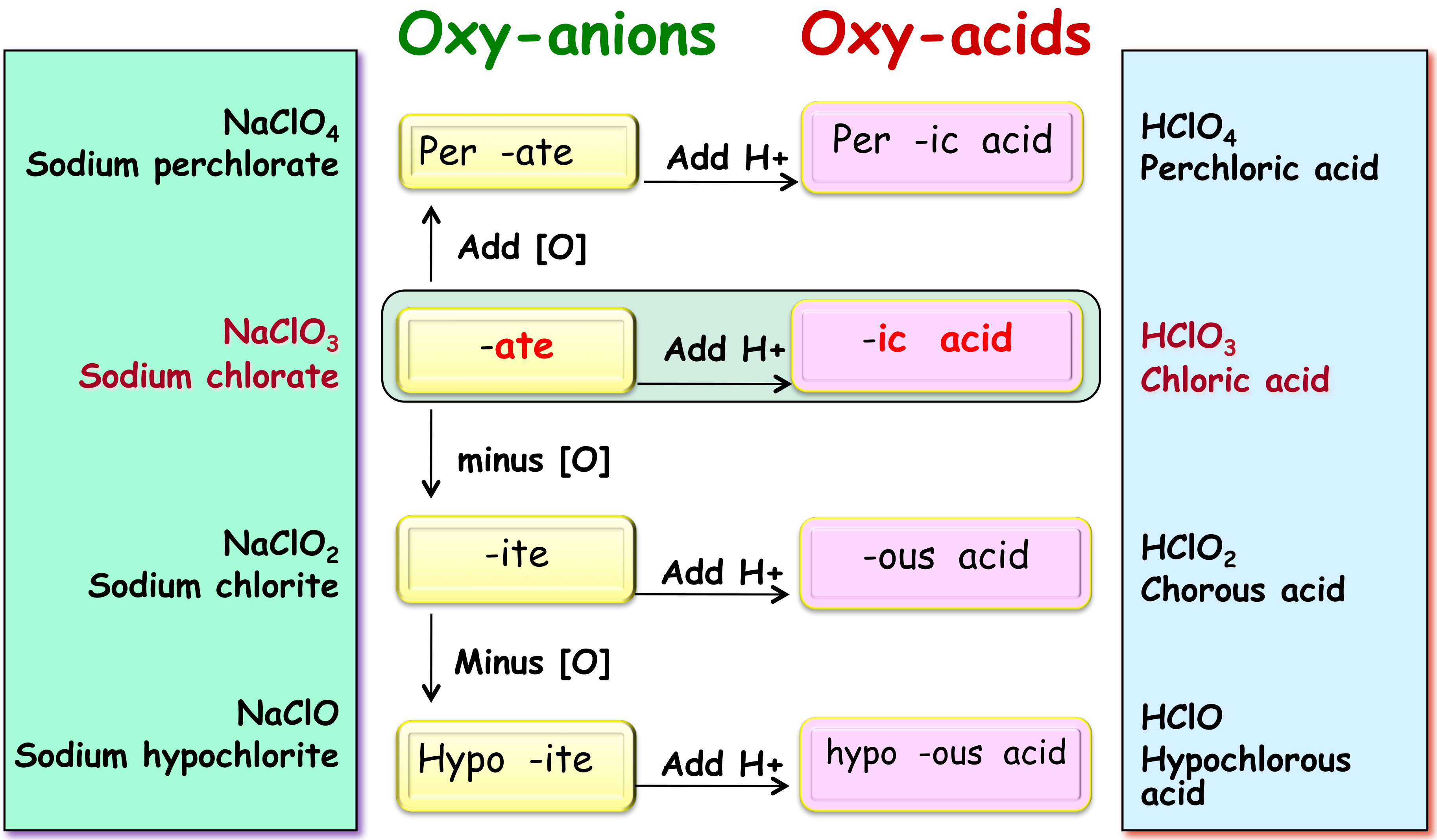
-3 -2 -1

Oxy-ions and oxy-Acids Nomenclature

Remember the -ate ions and start nomenclature from there.



Example of oxy ion salt and oxy acids.



... helpful hints about anions

-ide	-ate	-ite
elemental ions	oxy ions	oxy ions
X^{-m}	XO_n^{-m}	XO_{n-1}^{-m}
P^{-3} phosphide	PO_4^{-3} phosphate	PO_3^{-3} phosphite
As^{-3} arsenide	AsO_4^{-3} arsenate	AsO_3^{-3} arsenite
S^{-2} sulfide	SO_4^{-2} sulfate	SO_3^{-2} sulfite
Se^{-2} selenide	SeO_4^{-2} selenate	SeO_3^{-2} selenite
N^{-3} nitride	NO_3^- nitrate	NO_2^- nitrite
Cl^- chloride	ClO_3^- chlorate	ClO_2^- chlorite
Br^- bromide	BrO_3^- bromate	BrO_2^- bromite
I^- iodide	IO_3^- iodate	IO_2^- iodite

Family of oxyions

Group IV

Boron

BO_3^{-3} borate

Carbon

CO_3^{-2} carbonate

Group V

Nitrogen

NO_3^- nitrate

NO_2^- nitrite

Phosphorus

PO_4^{-3} phosphate

PO_3^{-3} phosphite

Arsenic

AsO_4^{-3} Arsenate

AsO_3^{-3} Arsenite

Group VI

Sulfur

SO_4^{-2} sulfate

SO_3^{-2} sulfite

Selenium

SeO_4^{-2} selenate

SeO_3^{-2} selenite

Tellurium

TeO_4^{-2} Tellurate

TeO_3^{-2} Tellurite

Group VII

Chlorine

ClO_4^- perchlorate

ClO_3^- chlorate

ClO_2^- chlorite

ClO^- hypochlorite

Bromine

BrO_4^- perbromate

BrO_3^- bromate

BrO_2^- bromite

BrO^- hypobromite

Iodine

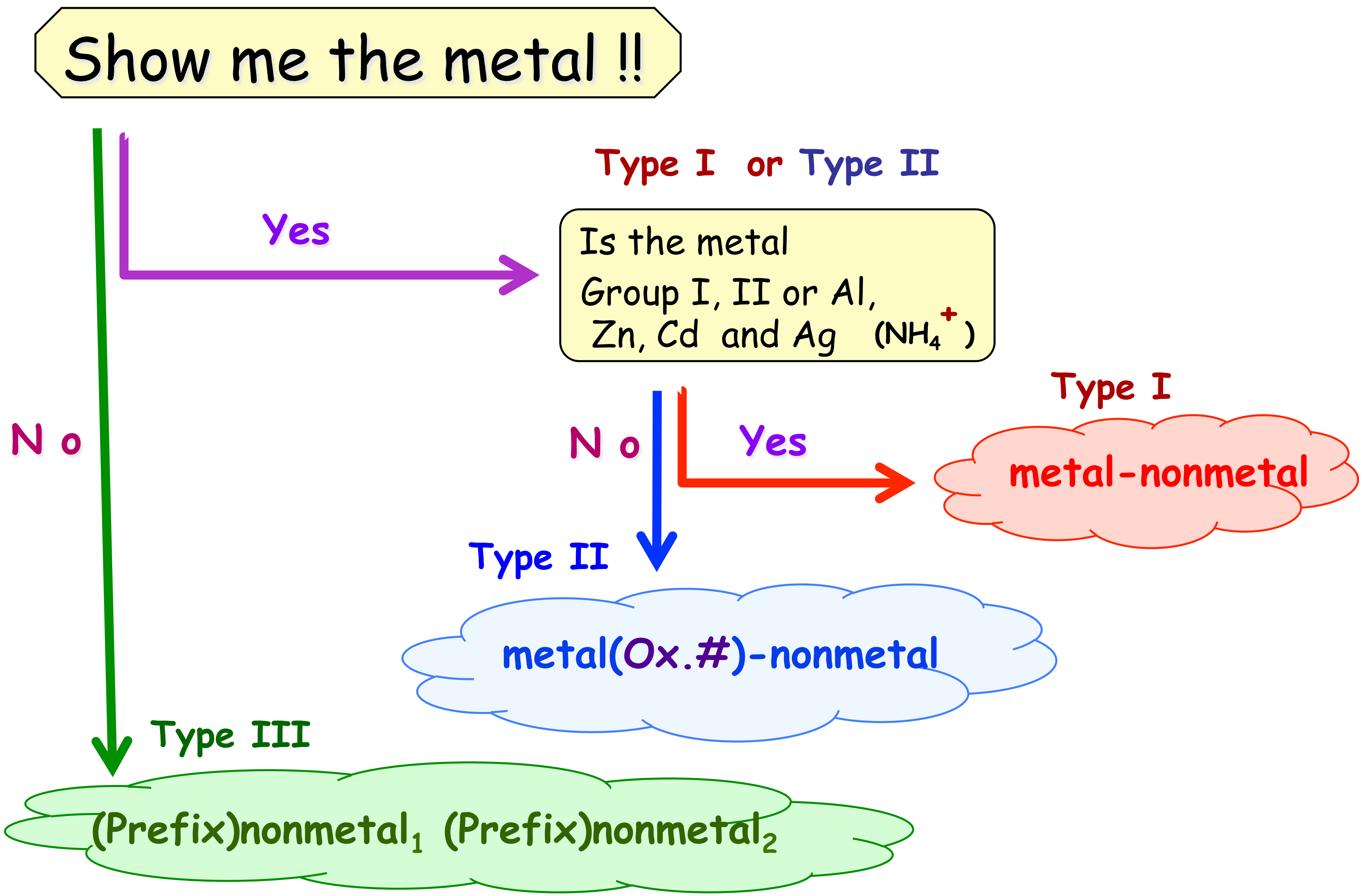
IO_4^- periodate

IO_3^- iodate

IO_2^- iodite

IO^- hypoiodite

Determining Type (I, II, III) of Compound



Type I, II and III

Type	Elemental anion	Polyatomic anion
I (Metal - nonMetal) Cation; Rep Metal Cation - Anion	**Cation - Anion(ide)** Al_2O_3; Aluminum oxide Ag_2S; Silver sulfide	**Cation - Anion** $Cd(NO_3)_2$: cadmium nitrate $(NH_3)_2SO_4$: ammonium sulfate
II (Metal - nonMetal) (Transition) metal Cations (Oxid #) – Anion **Old method (Fe^{+3} vs Fe^{+2})** higher ox.st. -ic lower ox. st. -ous	**Cation (oxidation state) Anion(ide)** $FeCl_3$; Iron(III) chloride PbS_2; Lead(IV) sulfide Cation(ic) - Anion(ide) $FeBr_3$; Ferric bromide Iron(III) bromide Cation(ous) – Anion(ide) Fe_3N_2 ; Ferrous nitride Iron(II) nitride	**Cation (oxidation state) Anion** $Sn(C_2H_3O_2)_2$; Tin(II) acetate Au_3PO_3; Gold(I) phosphite Cation(ic) – Polyatomic anion $Fe(NO_3)_3$; Ferric nitrate Iron(III) nitrate Cation(ous) – Polyatomic anion $Fe(NO_2)_2$; Ferrous nitrite Iron(II) nitrite
III Molecular compounds- Compounds which contains nonmetal (Prefix) nonmetal $_1$ - (Prefix) nonmetal $_2$	(Prefix) nonmetal $_1$ - (Prefix) nonmetal $_2$ (ide) • Prefixes are indication of the number of atoms: mono-, di-, tri-, tetra-, penta-, hexa- • order of naming nonmetal $_1$ & nonmetal $(ide)_2$ nonmetal $_1$ is to the left and bottom of nonmetal $_2$ based on it is named first in the nomenclature scheme. Si - C - As - P - N - H - Se - S - I - Br - Cl - O - F • S & 3 O forms SO_3 ; Sulfur trioxide • 2 P & 5 O forms P_2O_2 ; Diphosphorus pentaoxide	

Lewis Structure via Bond Table

1. (Connectivity) From the Chemical Formula, determine the atom connectivity for the structure.

   i. Given a chemical formula, AB_n, A is generally the central atom and B flanks the A atom. i.e., NH_3, NCl_3, NO_2. In these examples, N is central in each structure.

   ii. H and F are never central atoms even if their elemental symbol is first, i.e., H_2O.

2. (# of Bond) Determine the number of bonds in the compound, by calculating the theoretical Octet electrons (Oe) minus the total valence electrons (TVe).
   Oe is the theoretical number of electrons necessary for each atom in the structure to obtain a Noble Gas electron configuration, while TVe is the actual number of total valence electron for each atom in the structure.

3. (Remaining e-) Calculate the number of remaining electrons in the compound by taking the total valence electron (TVe) minus the number of electrons that was used to form bonds.

Complete Lewis structure by drawing atomic connectivity. Write bonds in the structure and the place remaining electrons to selected atoms in the structure to give each atom an octet. Keep in mind that the H-atom is satisfied with 2 electrons.

Lewis Dot Structure of ClO_4^- by Bonds Table

A. Calculate (Oe-) and (TVe)

ClO_4^-	Oe	Tve
1 Cl	1•(8)= 8	1•(7) = 7
4 O	4•(8)= 32	4•(6)= 24
Chg		1
	40	32

B. Number of Bonds.

$$\# \text{ bonds} = \frac{(40 - 32)}{2} = \frac{8}{2} = 4 \text{ bonds}$$

C. Remaining electrons.

Remaining e- = 32 – 8 = 24 e^- Remaining

Writing the Lewis Structure:

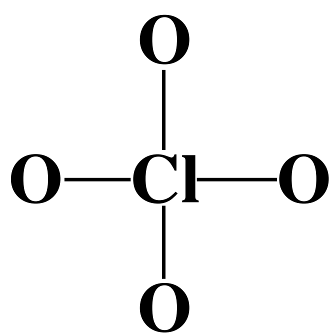
1. Write atom connectivity for ClO_4^-.

O

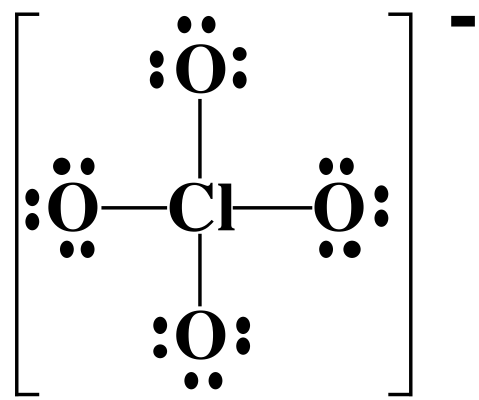
O Cl O

O

2. Draw the four bonds in the structure.



3. Place the remaining 24 electrons in the structure such that each atom has an octet to complete the Lewis Structure



Lewis Structure Summary

Compounds, elements comes together:

i) electrons are shared between elements

- •if there is mutually sharing, covalent compounds forms
- •if there is unequal sharing, polar covalent compounds forms.

ii) electron transfer occurs, ionic compounds forms (next section).

Lewis Structure Determination:

i) Molecular Formula

ii) Atomic Sequence (H and F are terminal)

iii) Determine the # of bonds

Oe^- and TVe^-

# of Bonds = $(Oe - TVe^-) / 2$

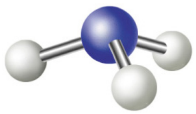
iv) Determine remaining electrons

$Re = (TVe^-) - (\# e^- \text{ in bonding})$

v) Make sure all atoms satisfy octet rule

(Except H which is satisfied with 2 electrons)

VSEPR Table

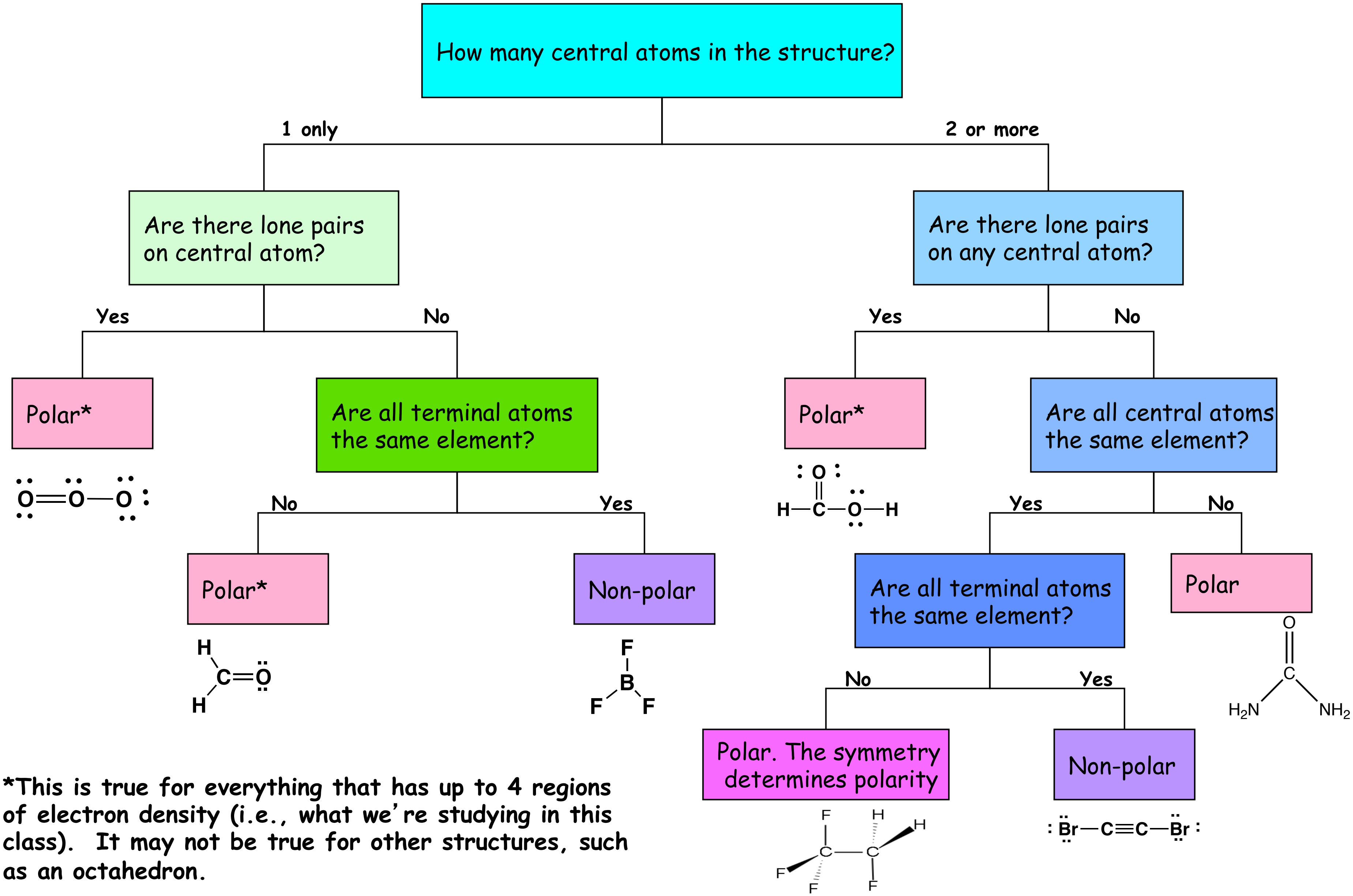


- **Molecular Geometry Determination Guide**

Valoence Shell Electron Pair Reuplsion (VSEPR) Table

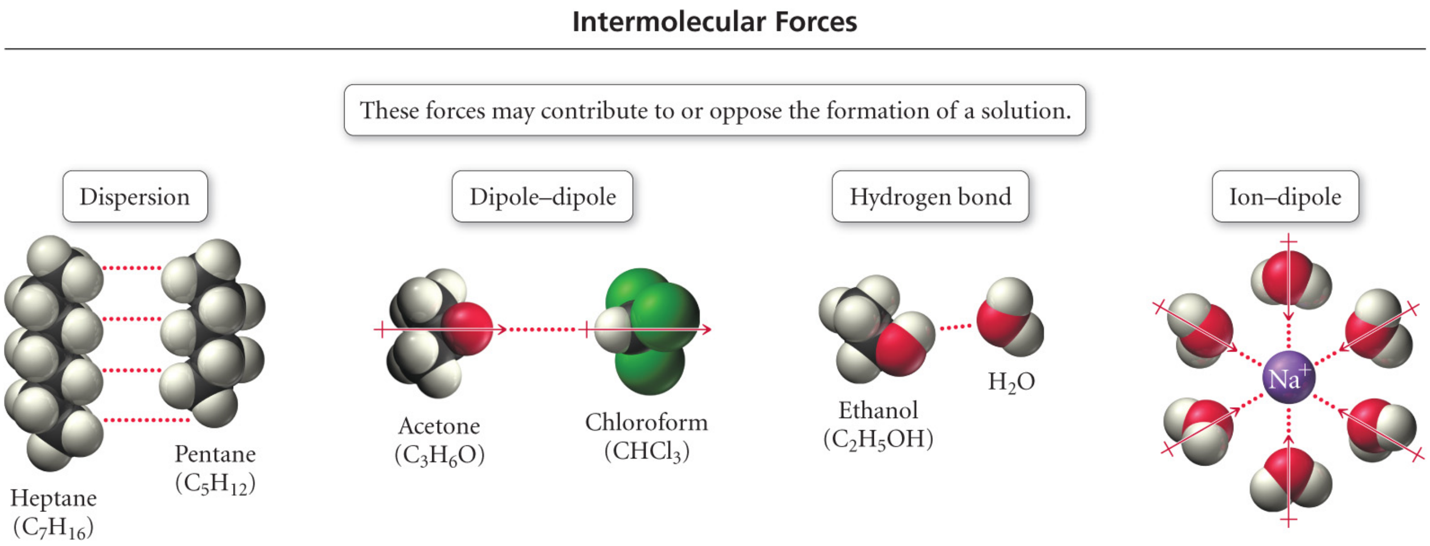
Electron Domains (Regions)	AE_n	Electronic Geometry	# Bonded Atoms (Coord #)	Lone pair on central atom	AB_mE_n	Molecular Geometry	Bond angle & Hybridization
2	AE_2	E A E Linear	2	0	AB_2	B—A—B Linear	180° sp
3	AE_3	Trigonal	3	0	AB_3	Trigonal	120° sp^2
			2	1	AB_2E	Bent	< 120° sp^2
4	AE_4	Tetrahedral	4	0	AB_4	Tetrahedral	109.5° sp^3
			3	1	AB_3E	Pyramidal	< 109.5° sp^3
			2	2	AB_2E_2	Bent	< 109.5° sp^3

Determining Molecular Polarity Flowchart



*This is true for everything that has up to 4 regions of electron density (i.e., what we're studying in this class). It may not be true for other structures, such as an octahedron.

Intermolecular Forces: Relative Strength



Interaction	Example	Energy
ion- ion	Na^+ Cl^-	400 - 4000 kJ
Covalent Bonds	H - H	150-1100 kJ
ion-dipole (I-D)	Na^+ HCl	40-600 kJ
dipole - dipole (D-D)	HCl HCl	5-25 kJ
dipole - induce dipole (D-ID)	HCl O_2	2-10 kJ
London Dispersion (LD)	N_2 N_2	0.05 - 40 kJ

- H-Bond (10- 40 kJ/mol

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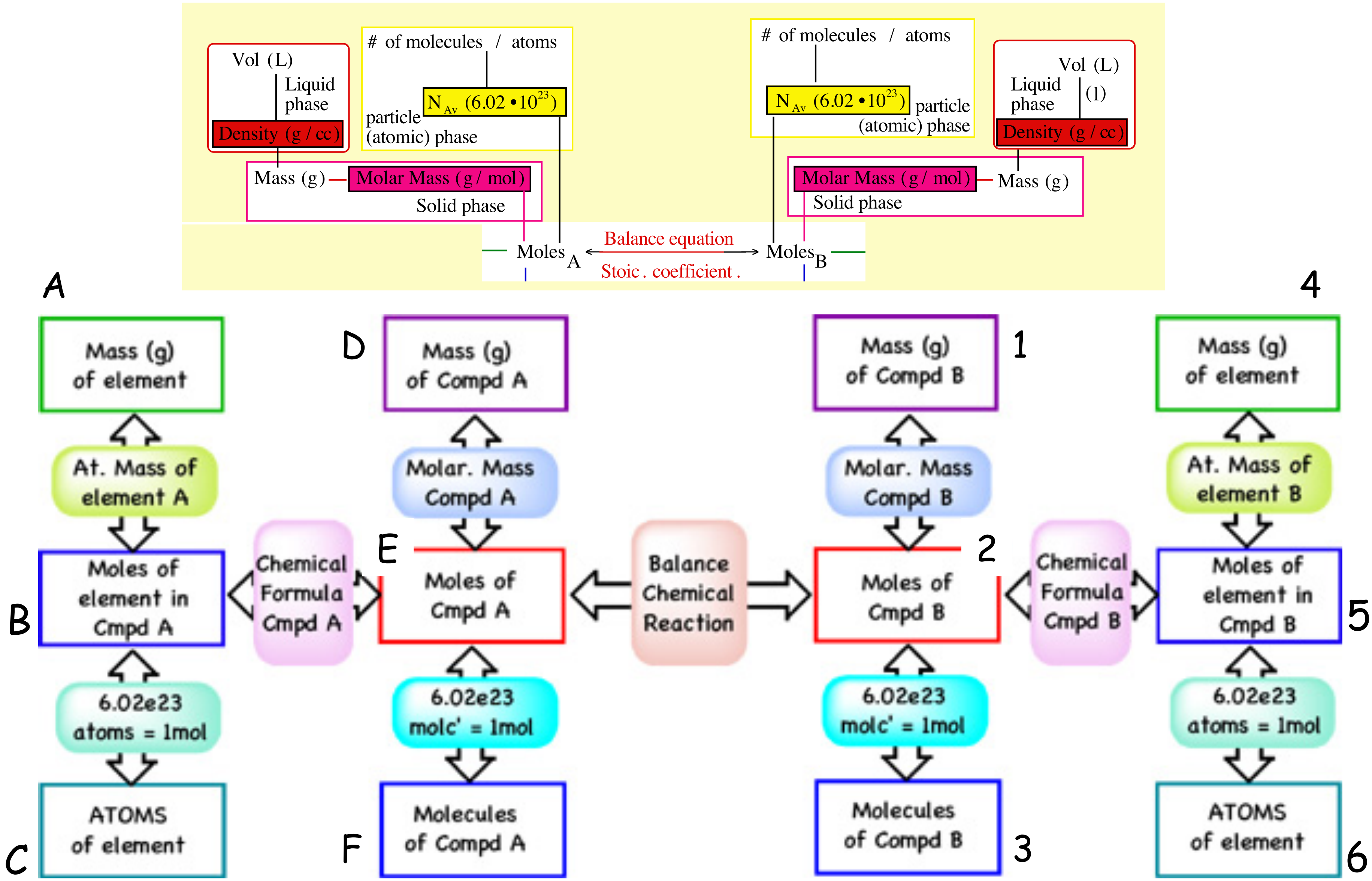
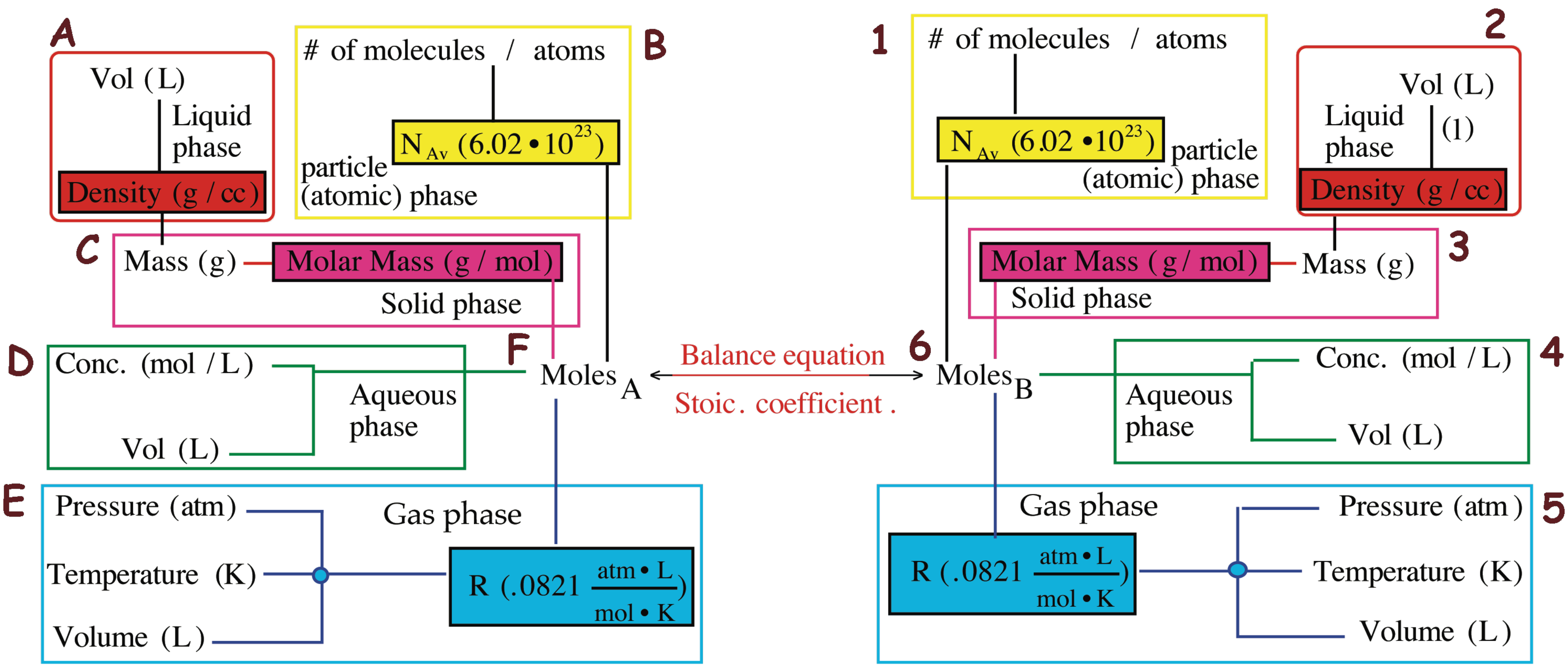
Chemical Reactions and Quantities
Chapter 6

Stoichiometry Map 1



Stoichiometry Map 2



Solubility Rules

Soluble Substances		Insoluble Substances	
Containing-	*Exceptions*	*Containing-*	*Exceptions*
Nitrates (NO_3^-) Perchlorates (ClO_4^-) Acetates (CH_3CO_2-)	None	Carbonates (CO_3^{2-}) Chromates (CrO_4^{2-}) Phosphates (PO_4^{3-}) Sulfides (S^{2-})	**Alkali and NH_4+**
Halogens (X-) Cl-, Br-, I-	**Ag, Hg & Pb.**	Hydroxides (OH^-)	**Ca, Ba, Sr, Alkali & NH_4+**
Sulfates (SO_4^{2-})	**Ca, Ba, Hg and Pb**	**Soluble - dissolve, no precipitate (aq -phase) insoluble (or slightly sol.) - does not dissolve, precipitate forms. (s-phase)**	
Alkali (Group1A) NH_4+	None		

Double Displacement Type Reaction: $Mg(NO_3)_{2(aq)} + 2KOH_{(aq)} \leftrightarrows$?

a) **Molecular equation**:

$$Mg(NO_3)_{2(aq)} + 2KOH_{(aq)} \leftrightarrows 2KNO_{3(aq)} + Mg(OH)_{2\ (s)}$$

b) **Complete ionic equation:**

$$Mg^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)} + 2\ K^+_{(aq)} + 2\ OH^-_{(aq)} \leftrightarrows 2\ K^+_{(aq)} + 2\ NO_3^-{}_{(aq)} + Mg(OH)_{2\ (s)}$$

c) **Net ionic equation**:

$$Mg^{2+}_{(aq)}) + 2\ OH^-_{(aq)} \leftrightarrows Mg(OH)_{2\ (s)}$$

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Solutions
Chapter 8

Concentrations and Dilutions

•Ways of expressing concentration-

-**Molarity(M)** - moles solute / Liter solution

-**Concentration by parts**

-Mass percent, % m (m:m)- (grams solute/Total grams of solution) *100 → pph

-Mass-Vol percent, (m:v)- (grams solute/ Total ml volume solution) *100→ pph

-Vol-vol percent, (v:v)- (volume ml solute/ Total ml volume solution) *100→ pph

Note that if the multiplier is $1 \cdot 10^6$ instead of 100, then the unit is ppm.

ppm (m:m) = (grams solute/Total grams of solution) x $1 \cdot 10^6$ → ppm

If the multiplier is 109, then the unit is ppb.

ppb (m:m) = (grams solute/Total grams of solution) x $1 \cdot 10^9$ → ppb

-**Dilution Equation:** $C_1 \cdot V_1 = C_2 \cdot V_2$

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Acid and Bases
Chapter 10

Determining pH, pOH, $[OH^-]$, $[H_3O^+]$

Use this chart to determine acid and base concentration at (25°C)

$K_w = 1.0 \cdot 10^{-14}$ & $pK_w = 14.00$

$[H_3O^+] \cdot [OH^-] = K_w$ & $pH + pOH = pK_w$