<u>Unit 1: Foundations</u>

Chapter Chemical Foundations & Atoms, Molecules etc. (Chemistry Honors)

Notes #1 <u>Ch. Chemical Foundations</u>

<u>I. Significant Figures</u>

-are digits in a number that have been measured.

<u>Rule:</u>

Every digit is significant, <u>except</u>:

-leading zeros in small numbers.

-trailing zeros in numbers <u>without</u> a decimal point.

In 0.0023 only the 2 & 3 count, in 5000 only the 5 counts.

5000**.** (all count)

Ex. 1) State the number of significant figures.

4.23	3	
5.20	3	
0.000059	2	(leading zeros)
5.09	3	
20.090	5	
320.00	5	
10200	3	(trailing zeros, no decimal point)

<u>II. Exponential Notation</u>

-is a way of writing numbers using <u>only significant figures</u>.

#.## X 10^{power} (1 digit, then decimal, then rest of number)

256 = 2.56 X 100 = **2.56 $X10^{2}$**

43.9 = **4.39 $X10^{1}$**

0.02 = **2. $X10^{-2}$** Big #, positive power!

0.000490 = **4.90 $X10^{-4}$** Small #, negative power!

10200 = **1.02 $X10^{4}$** *<u>non significant figures are not included</u>

<u>III. Rules for Rounding with Significant Figures</u>

<u>a) Addition/Subtraction</u>

The answer is rounded to the decimal place of the least accurate digit in the problem.

Ex. 1) 429.3 tenths *tenths is least accurate (fewest decimal places)

37.45 hundredths

<u>+ 1.93 hundredths</u>

468.68 round it to tenths

468.7 = **4.687 $X10^{2}$**

Ex. 2) $1.956 \times 10^2 - 2.3 \times 10^1$

195.6	tenths	
-23.	ones	*ones is least accurate
172.6	round to ones	
173 = **1.73×10^2**		

**Take numbers out of exponential notation.

b) Multiplication/Division

The answer should have the same number of significant figures as the term with the least number of significant figures in the problem.

Ex. 1) $4.2 \times 10^4 \times 2.43 \times 10^2 = (4.2 \times 10^4) \times (2.43 \times 10^2) = 10206000 =$ **1.0×10^7**

1st number has 2 sig fig, 2nd number has 3 sig fig, 2 is least

Ex. 2) $4.9 \times 10^{-2} \times 3.11 \times 10^2 / 3.97 \times 10^3 = 0.003838539 =$ **3.8×10^{-3}**

1st number has 2 sig fig, 2nd number has 3 sig fig, 3rd number has 3 sig fig, 2 is least

$X10^n$ (-)

On calculator: 4.9 **EXP or EE** **+/-** 2 **X** 3.11 **EXP or EE** 2 ÷ 3.97 **EXP or EE** 3 =

Some new calculators have a **$X10^n$ button, instead of **EXP** or **EE** . (Do not confuse this with the **10^x** button, which is for inverse log! Never type in **X 1 0** separately!)

Some new calculators have a **(-) or a $+ \rightarrow\leftarrow$ **-** button, instead of a **+/-** button.

On the calculator, these buttons are usually on top of the 7 or 8 or by the equals (=) sign.

Exact Numbers

-have an infinite number of significant figures.

4 eggs = 4.00000000

Averaging 3 numbers $\frac{\# + \# + \#}{3}$ the 3 is exact, 3.00000000

These exact numbers will not limit you, look at the other numbers in the problem for rounding.

Notes #2 IV. Combined Operations

Steps:
1) Work out the problem on the calculator.
2) Then go back and find what the numbers should be rounded to.
3) Round off the original answer.

Ex. 1) (4.23 + 5.6) (3.13 + 4.937) = (9.83) (8.067) = 79.2986

4.23
+5.6 ←tenths least accurate
9.8 rounded to tenths

3.13 ←hundredths least accurate
+4.937
8.07 rounded to hundredths

(9.8) (8.07) = 79 = 7.9 X10^{1} rounded to 2 sig fig

↑

2 sig fig is the least amount

Ex. 2) $\frac{(1.53 + 2.961 + 37.0)}{(42.3 - 29.345 - 8.21)} = \frac{41.491}{4.745} = 8.7441517$

1.53
+2.961
+37.0 ←tenths least accurate
41.5 rounded to tenths

42.3 ←tenths least accurate
-29.345
-8.21
4.7 rounded to tenths

(41.5) / (4.7) = **8.7 = 8.7 X10^{0}** rounded to 2 sig fig

↑

2 sig fig is the least amount

Ex. 3) $\frac{(79.12 - 16.007 + 0.1)}{(49.30 + 24.970)} = \frac{63.213}{74.270} = 0.8511242$

79.12
-16.007
+0.1 ←tenths least accurate
63.2 rounded to tenths

49.30 ←hundredths least accurate
+24.970
74.27 rounded to hundredths

(63.2) / (74.27) = **0.851 = 8.51 X10^{-1}** rounded to 3 sig fig

↑

3 sig fig is the least amount

V. SI Units

Mass = grams (really kilograms)
Length = meters
Time = seconds
Volume = liters

Multipliers	Example with meters: **(**Same for grams, seconds & liters!)**
Mega (M)	1 Mm = 1 X10^{6} m
**kilo (k)	1 km = 1 X10^{3} m
**deci (d)	1dm = 1 X10^{-1} m
**centi (c)	1cm = 1 X10^{-2} m
**milli (m)	1 mm = 1 X10^{-3} m
micro (µ)	1 µm = 1 X10^{-6} m
nano (n)	1 nm = 1 X10^{-9} m
pico (p)	1 pm = 1 X10^{-12} m

length

1 m = 39.37 in
2.54 cm = 1 in
1 km = 0.621 mile
1 mile = 5280 ft

Mass

1kg = 2.205 lb
1 lb = 16 oz

Volume

1 L = 1.06 qt   ** 1 L = $1dm^3$ and 1 ml = $1cm^3$
1 gal = 3.773 L
1 gal = 4 qt

Notes #3 VI. Conversions

Ex. 1) Convert 2.00 hr to min

$$\frac{2.00\ \text{hr}}{} \Big| \frac{60\ \text{min}}{1\ \text{hr}} = 120\ \text{min} = \mathbf{1.20\ X10^{2}\ min}$$

1 hr = 60 min

Ex. 2) Convert 3.0 mi to km

$$\frac{3.0\ \text{mi}}{} \Big| \frac{1\text{km}}{0.621\ \text{mi}} = \mathbf{4.8\ X\ 10^{0}\ km}$$

1 km = 0.621 mi

Ex. 3) Convert 49.6 in to km

in → m → km {can go through other combinations of units}

$$\frac{49.6\ \text{in}}{} \Big| \frac{1\text{m}}{39.37\ \text{in}} \Big| \frac{1\text{km}}{1\ X10^{3}\ \text{m}} = \mathbf{1.26\ X10^{-3}\ km}$$

1 m = 39.37 in, 1km = 1 X10^3 m

Ex. 4) Convert 34.6 mi to mm

mi → km → m → mm

$$\frac{34.6\ \text{mi}}{} \Big| \frac{1\ \text{km}}{0.621\ \text{mi}} \Big| \frac{1\ X10^{3}\ \text{m}}{1\ \text{km}} \Big| \frac{1\text{mm}}{1\ X10^{-3}\ \text{m}} = \mathbf{5.57\ X10^{7}\ mm}$$

1 km = 0.621 mi, 1km = 1 X10^3 m, 1mm = 1 X10^{-3} m

Ex. 5) Convert 15.6 kg/m^3 to g/cm^3

$$\frac{15.6\text{ kg}}{\text{m}^3}\left|\frac{1\text{ X}10^3\text{ g}}{1\text{ kg}}\right|\frac{1\text{ X}10^{-6}\text{ m}^3}{1\text{cm}^3} = \mathbf{1.56\ X10^{-2}\ g/cm^3}$$

$$1\text{ kg} = 1\text{ X}10^3\text{ g}, \quad (1\text{cm})^3 = (1\text{ X}10^{-2}\text{ m})^3$$
$$1\text{cm}^3 = 1\text{ X}10^{-6}\text{ m}^3$$

Ex. 6) Convert 1.34 X10^6 cm/sec to mi/day

$$\frac{1.34\text{ X}10^6\text{ cm}}{\text{sec}}\left|\frac{1\text{ X}10^{-2}\text{ m}}{1\text{cm}}\right|\frac{1\text{km}}{1\text{ X}10^3\text{ m}}\left|\frac{0.621\text{ mi}}{1\text{ km}}\right|\frac{60\text{ sec}}{1\text{ min}}\left|\frac{60\text{ min}}{1\text{hr}}\right|\frac{24\text{ hr}}{1\text{ day}} = \mathbf{7.19\ X10^5\ mi/day}$$

Notes #4 VII. Temperature & Density

A) Temperature Conversions

$^oC = 5/9\ (^oF - 32)$ $K = {^oC} + 273$

^oC = Celsius ^oF = Fahrenheit K = Kelvin (0 K = Absolute Zero)
-lowest possible temperature,
all motion stops, no kinetic energy.

Ex. 1) What is 154 K in ^oF?

$K = {^oC} + 273$
$154\ K = {^oC} + 273$
$-119 = {^oC}$

$^oC = 5/9\ (^oF - 32)$
$-119\ ^oC = 5/9\ (^oF - 32)$
$9/5\ (-119\ ^oC) = {^oF} - 32$
$-214.2 = {^oF} - 32$
$-182 = {^oF}$

Ex.2) What is 72 ^oF in Kelvin?

$^oC = 5/9\ (^oF - 32)$
$^oC = 5/9\ (72\ ^oF - 32)$
$^oC = 22$

$K = {^oC} + 273$
$K = 22\ ^oC + 273$
K = 295

B) Density

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \frac{\text{(grams)}}{(\text{cm}^3)}$$

mass is constant, weight changes (weight = mass X gravity)

Density of water = 1 g/cm^3
O_2 = 1.33 X 10^{-3} g/cm^3
Au = 19.32 g/cm^3 L. aurum
Al = 2.70 g/cm^3
Ag = 10.5 g/cm^3 L. argentum

Ex. 1) Bismuth has a density of 9.80 g/cm^3. What is the mass of 4.32 ml of Bi?

$$\frac{4.32\ \text{ml}}{} \left| \frac{1\ \text{cm}^3}{1\ \text{ml}} \right. = 4.32\ \text{cm}^3$$

D = m/v $\quad 9.80\ \text{g/cm}^3 = \dfrac{m}{4.32\ \text{cm}^3}$

4.23×10^1 g = m

Ex. 2) Iron has a density of 7.87 g/cm^3. What volume would 2.46 X10^{-2} kg of Fe occupy?

$$\frac{2.46 \times 10^{-2}\ \text{kg}}{} \left| \frac{1 \times 10^3\ \text{g}}{1\ \text{kg}} \right. = 24.6\ \text{g}$$

D = m/v $\quad 7.87\ \text{g/cm}^3 = \dfrac{24.6\ \text{g}}{V}$

7.87 (V) = 24.6 **When in doubt, cross multiply!

V = 24.6 / 7.87

v = 3.13 X10^0 cm^3

#5 Notes

VIII. States of Matter**

Matter is anything that occupies space and has mass.

1) Solid: particles arranged in a rigid pattern, definite shape & volume.

2) Liquid: particles close together in no pattern, takes the shape of its container, definite volume.

3) Gas: no definite shape or volume (shape and volume of container)

4) Plasma: hot gas-like mixture over 5000 ^oC, collisions cause some electrons to be knocked off, creating (+) ions. This charged mixture that conducts electricity is plasma.

IX. Mixtures

physical

[Heterogeneous Mixtures] ↔ **[Homogeneous Mixture = Solutions]**

Parts are visible *methods* parts are indistinguishable

↕ *physical methods*

Pure Substances (Homo)

↓ ↓

Elements (Homo) ↔ **Compounds (Homo)**

(One type of atom) (2 or more types of atoms)

chemical methods

Ex. 1) Heterogeneous or Homogeneous?

a) salt poured into water	**Hetero**
b) salt is stirred/mixed into water	**Homo**
c) soda	**Hetero**
d) flat soda	**Homo**
e) unmixed Kool Aid in water	**Hetero**
f) apple	**Hetero**

Ex. 2) Element or Compound? (=Pure Substance = Homo)

a) salt	**compound**
b) isopropyl alcohol	**compound** *look up formula in textbook or on internet
c) gold	**element**
d) oxygen	**element**
e) sugar	**compound**

X. Physical/Chemical Characteristics/Changes**

Physical: melting, freezing, boiling, solubility, changing shape, malleability, conductivity.
Chemical: burning, exploding, reacting with acid, toxicity.

Ch. Atoms, Molecules, Ions

I. Law of Conservation of Mass** (Lavoisier):
Mass cannot be created or destroyed.

II. Law of Definite Proportion** (Proust):
A given compound always contains exactly the same proportion of elements by mass.

III. Law of Multiple Proportions** (Dalton):
When 2 elements form more than 1 compound,
for a fixed mass of one element,
the masses of the second element are related to each other by small whole numbers (1, 2, 3 etc.)

	mass of Fe	mass of Cl	
$FeCl_2$	56 g	70 g	
$FeCl_3$	56 g	105 g	105:70 → 3:2
	(fixed)		**Subscripts must be whole numbers!

H_2O, H_2O_2

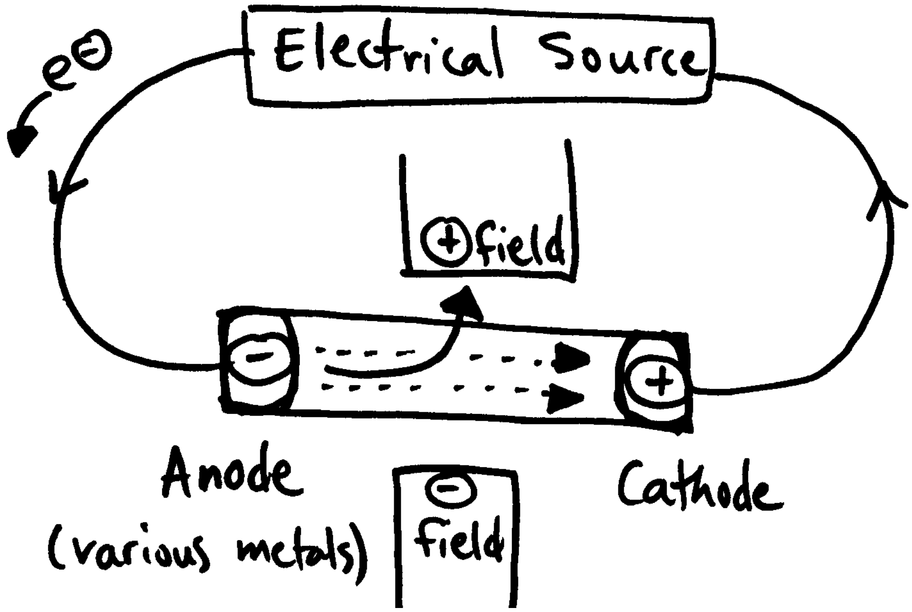
IV. Dalton's Atomic Theory**:
A) Each element is made up of tiny particles called atoms.
B) The atoms of a given element are identical; the atoms of different elements are different.
C) A compound always has the same ratio and types of atoms.
(like Law of Definite Proportions)
D) Chemical reactions involve reorganization of the atoms to form new compounds (the atoms themselves are unchanged).
(like Law of Conservation of Mass)

#6 Notes

V. Avogadro's Hypothesis**:
At the same temperature and pressure, equal volumes of different gases contain the same number of particles. (22.4 L = 1 mol = 6.022×10^{23} molecules of the gas)

VI. Atomic Models

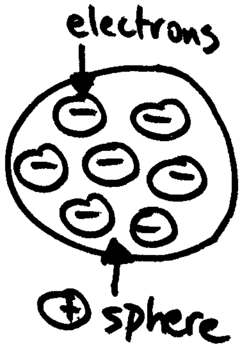
J.J. Thompson and the Cathode Ray Tube



When a high voltage is applied, a cathode ray is produced. The particles he called electrons were (-), since they went toward an applied (+) electrical field and were repelled by an applied (-) electrical field.
**Since electrons were produced from electrodes of various metals, he postulated that all atoms must contain (-) electrons.

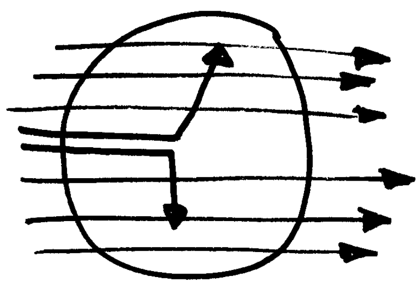
Millikan's Oil Drop Experiment
A falling oil drop can be halted by adjusting a voltage across 2 plates. With the voltage and the mass of an e^-, he could calculate the charge on a single e^-. (The charge on an oil drop is always a whole number multiple of the charge on a single e^-, so an e^- is not divisible into other particles.)

A) Plum Pudding Model (J.J. Thompson)
All atoms contain (-) electrons. Since atoms are neutral, there must also be (+) somewhere.



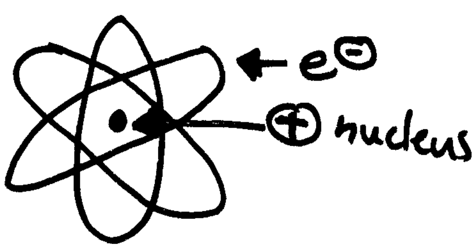
**(-) electrons are randomly scattered in a spherical cloud of (+) charge.

B) Rutherford's Gold Foil Experiment
Radioactivity was discovered by Becquerel in 1896. (α - particles are helium atoms)



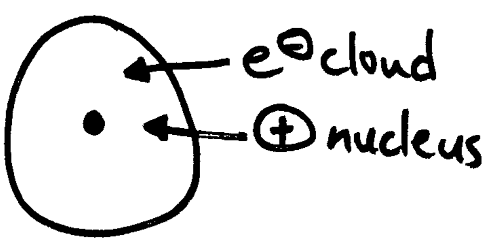
When α – particles are shot at an Au atom, the center repels the α – particles, so there must be a concentration of charge in the center.

Rutherford Model**



**A (+) nucleus is surrounded by orbiting (-) electrons.

C) Modern Model **



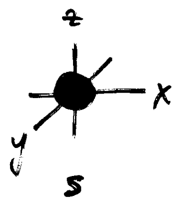
**A (+) nucleus (diameter 10^{-13} cm), containing (+) protons and neutral neutrons is surrounded by an electron cloud (diameter 10^{-8} cm).

{Tennis ball nucleus 4.5 cm, e- cloud out 4.5 km ≈ 2.8 mi}

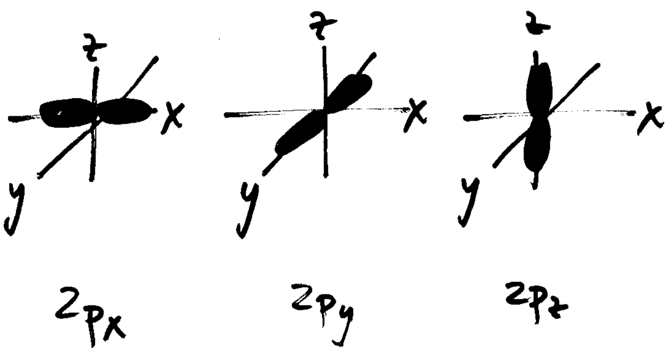
**electrons are in layers of many differently shaped orbitals (s, p, d, f)

Electron Orbital Diagrams

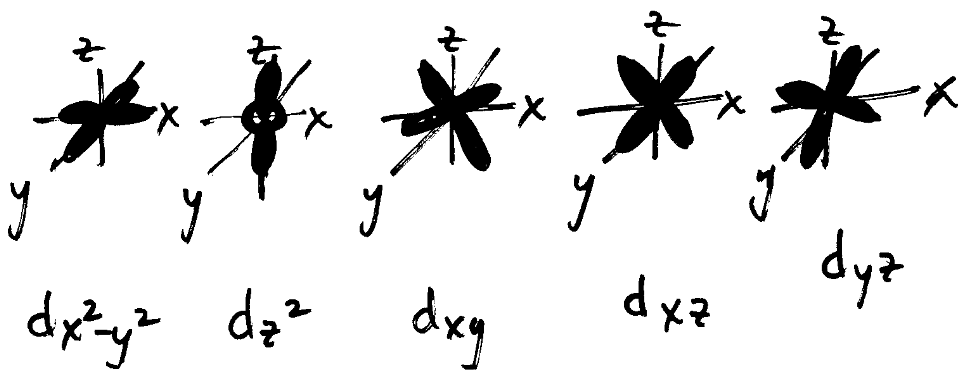
I. s – orbitals (1 type)



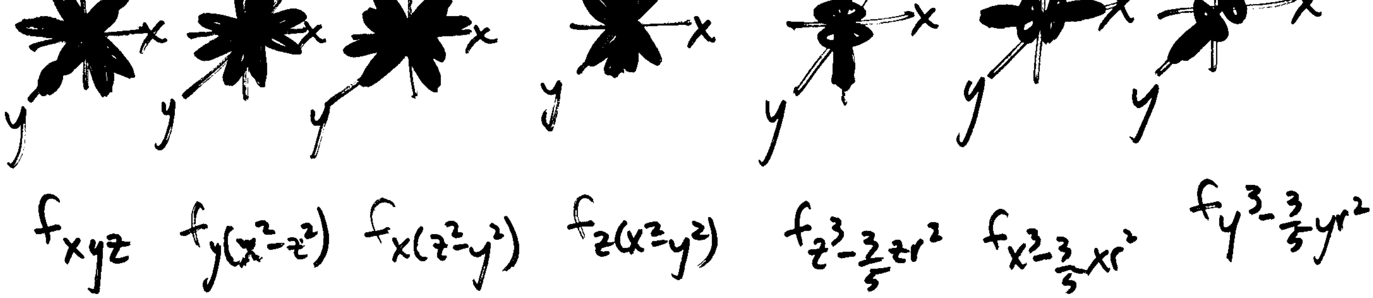
II. p – orbitals (3 types)



III. d – orbitals (5 types)



IV. f – orbitals (7 types)



Isotopes** are atoms of the same element, but with different mass. They contain different amounts of neutrons.

	Mass	Charge
Electron	9.109×10^{-31} kg	-1
Proton	1.672×10^{-27} kg	+1
Neutron	1.675×10^{-27} kg	neutral

Top # = mass # → ^{23}Na
Bottom # = atomic # → 11

Atomic # = # of Protons = # of Electrons
Mass # = # of Protons + # of Neutrons {To find mass}
Mass # – Atomic # = # of Neutrons {To find neutrons}

Given:	neutrons	protons	electrons	charge
Ca = $^{40}_{20}Ca$ (find neutrons)	40-20 = 20	20	20	0
^{235}U = $^{235}_{92}U$	235 – 92 = 143	92	92	0

Notes #7

$Ca^{2+} = {}^{40}_{20}Ca^{2+}$	40-20= 20	20	20-2=18	2+	(+) charge = lose electrons
$F^{1-} = {}^{19}_{9}F^{1-}$	19-9 = 10	9	9+1= 10	1-	(-) charge = gain electrons

Given: 26 protons, 30 neutrons, +3 charge.

26 + 30 = ${}^{56}_{26}Fe^{+3}$ (find mass)	30	26	26 -3 = 23	3+	(+) charge = lose electrons

VII. Periodic Table**

A) Metals: left side of table
- Conduct heat and electricity
- Malleable, ductile
- Lustrous
- Lose e^- to form (+) ions

B) Nonmetals: right side of table
- Gases or brittle solids
- Poor conductors
- Gain e^- to form (-) ions

C) Metalloids (Semi-metals, Semiconductors):
- Properties of metals and nonmetals
- B, Si, Ge, As, Sb, Te, Po, At

See Periodic Table with Group Names and label groups/families:
Alkali Metals, Alkaline Earth Metals, Transition Metals, Halogens, Noble Gases, Lanthanides, Actinides, Metals, Nonmetals, Metalloids.

VIII. Bonds

Chemical Bonds are the force that holds atoms together.

Covalent Bonds: are when electrons are shared.
(molecules have covalent bonds)

Ionic Bonds: are when electrons are transferred.
(ionic solids/ salts have ionic bonds)

NaCl

Loses 1 e^- ↓ ↓ Gains 1 e^-

Na^+ Cl^- → NaCl Na's e^- is transferred to the Cl.

Cations (+) Anions (-)

Polyatomic Ions: are covalently bonded atoms with a charge

NH_4^{1+}, SO_4^{2-}

IX. Writing Formulas for Ionic Compounds

The ions charges must (be multiplied to) make neutral compounds.

Ex. 1) sodium nitrate Na^{1+} NO_3^{-1} → **$NaNO_3$**

Ex.2) barium sulfate Ba^{2+} SO_4^{2-} → **$BaSO_4$**

Ex. 3) magnesium chloride Mg^{2+} Cl^{1-} → **$MgCl_2$**

X1 X2

+2 -2 = 0

Ex. 4) iron III hydroxide Fe^{3+} OH^{-1} → **$Fe(OH)_3$**

X1 X3

+3 -3 = 0

Ex. 5) lithium phosphate Li^{1+} PO_4^{3-} → **Li_3PO_4**

X3 X1

+3 -3 = 0

Ex. 6) cobalt III sulfite Co^{3+} SO_3^{2-} → **$Co_2(SO_3)_3$**

X2 X3

+6 -6 = 0

Ex. 7) chromium II phosphate Cr^{2+} PO_4^{3-} → **$Cr_3(PO_4)_2$**

X3 X2

+6 -6 = 0

Notes #8 X. Writing Formulas for Covalent Molecules

These compounds contain 2 nonmetals (both (-)), so they will share electrons in different combinations.
Prefixes: mono (1), di (2), tri (3), tetra (4), penta (5), hexa (6), hepta (7), octa (8), nano (9), deca (10).   **mono only used on 2nd element

Ex. 1) sulfur tetrafluoride S^{2-} F^{1-} → **SF_4**

Ex. 2) dinitrogen tetroxide N^{3-} O^{2-} → **N_2O_4**

Ex.3) carbon monoxide $C^{?}$ O^{2-} → **CO**

Ex. 4) phosphorus trichloride P^{3-} Cl^{1-} → **PCl_3**

Ex. 5) sulfur hexafluoride S^{2-} F^{1-} → **SF_6**

XI. Naming Ionic Compounds

Ex. 1) $Sr(OH)_2$ **strontium hydroxide**

Ex. 2) $Cu(ClO_2)_2$ **copper II chlorite**
ClO_2^{1-} X 2 = -2, so Cu must be +2

Ex. 3) $CrBr_3$ **chromium III bromide**
Br^{1-} X 3 = -3, so Cr must be +3

Ex. 4) $Fe_2(SO_4)_3$ **iron III sulfate**
SO_4^{2-} X 3 = -6, so Fe side must be +6,
But there are 2 Fe, so each must be +3

Ex. 5) Cr_2O_3 **chromium III oxide**
O^{2-} X 3 = -6, so Cr side must be +6,
But there are 2 Cr, so each must be +3

Ex. 6) $Co_3(PO_4)_2$ **cobalt II phosphate**
PO_4^{3-} X 2 = -6, so Co side must be +6,
But there are 3 Co, so each must be +2

Ex. 7) $Ba(NO_3)_2$ **barium nitrate**

XII. Naming Covalent Molecules

Ex. 1) NO_2 N^{3-} O^{2-} prefixes: **nitrogen dioxide**
(mono never used on 1st element)
**Only the 2nd name gets the "-ide".

Ex. 2) ICl_3 I^{1-} Cl^{1-} **iodine trichloride**

**Think of metalloids as negative.
Metalloids (-) with Nonmetals (-) are covalent: $SiCl_4$ silicon tetrachloride
Metalloids (-) with Metals (+) are ionic: Fe_3As_2 iron II arsenide

XIII. Mixed Examples

Ex. 1) $K_2Cr_2O_7$ **potassium dichromate**
$Cr_2O_7{}^{2-}$

Ex. 2) PF_5 P^{3-} F^{1-} **phosphorus pentafluoride**

Ex. 3) $Pb(SO_4)_2$ **lead IV sulfate**
$SO_4{}^{2-}$ X 2 = -4, so Pb must be +4

Ex. 4) N_2O N^{3-} O^{2-} **dinitrogen monoxide**

Ex.5) BeF_2 **beryllium fluoride**

**Acids start with "H", see ion sheet.
If "H" used as (-), like NaH = sodium **hydride**

End of Notes (Assignments #9-10 are Review Assignments. There are no notes for these assignments.)