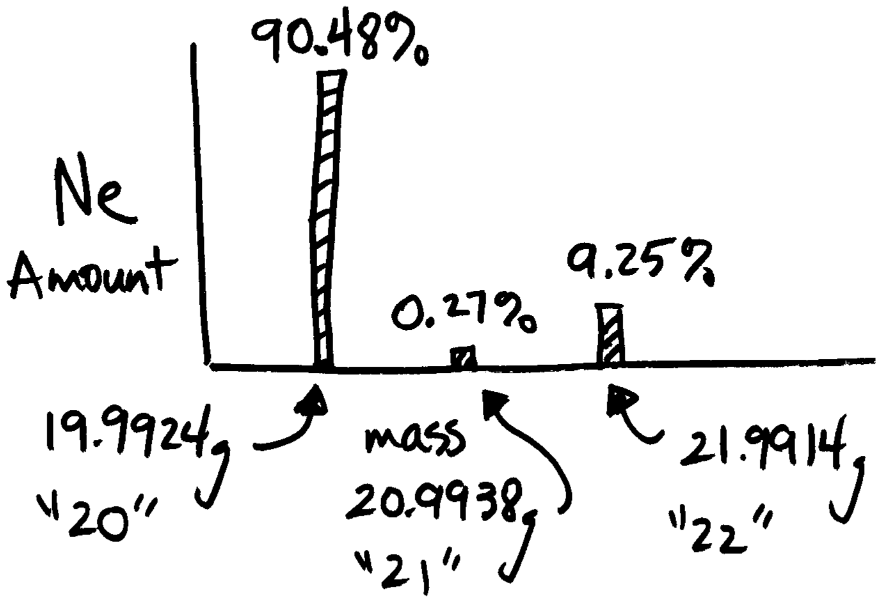
#11 Notes **Unit 2: Stoichiometry**

Ch. Stoichiometry

I. Atomic Mass

Atomic mass on the periodic table is the average mass of the naturally occurring isotopes weighted by their natural abundance. (This includes electrons and is not rounded to whole numbers, like the atomic mass # in the last chapter.)

Mass Spectrometers show the relative abundance by mass.



Ex. 1) Calculate the average atomic mass of Ne.

(% as decimal)(mass) + (% as decimal) (mass) + ….

(0.9048) (19.9924 g) + (0.0027) (20.9938 g) + (0.0925) (21.9914 g) =

18.09 + 0.05668 + 2.034 = **20.18 g**

**** or 20.18 amu (atomic mass units**)

Ex.2) Set up the equation for:

Find the % abundance for ^{35}Cl and ^{37}Cl, if the average mass is 35.453 amu. The mass of ^{35}Cl is 35.002 amu and ^{37}Cl is 36.998 amu.

(X) (35.002 amu) + (1- X) (36.998 amu) = 35.453 amu

II. The Mole (Mol)

Molar mass is the mass of 1 mol of a compound.

Ex.1) Find the molar mass of $Ca_3(PO_4)_2$.

3 Ca = 3 (40.078 g)
2 P = 2 (30.97376 g)
8 O = 8 (15.999 g)
310.174 g in 1 mol of $Ca_3(PO_4)_2$ = **310.174 g/mol**

1 mol = 6.022×10^{23} atoms or molecules

$$\begin{array}{cc} \text{atoms} & \text{molecules} \\ \downarrow & \downarrow \\ \text{Elements} & \text{Compounds} \end{array}$$

Ex. 2) Convert 59.32 g $Ca_3(PO_4)_2$ to molecules.

$$\frac{59.32 \text{ g } Ca_3(PO_4)_2}{} \left| \frac{1 \text{ mol}}{310.174 \text{ g}} \right| \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{1.152 \times 10^{23} \text{ molecules } Ca_3(PO_4)_2}$$

Ex. 3) Convert 4.59×10^{24} molecules $MgCl_2$ to grams.

$$\frac{4.59 \times 10^{24} \text{ molecules } MgCl_2}{} \left| \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \right| \frac{95.211 \text{ g}}{\underset{\substack{\uparrow \\ 1 \text{ Mg} + 2 \text{ Cl}}}{1 \text{ mol}}} = \mathbf{726 \text{ g } MgCl_2}$$

Ex. 4) Convert 45.3 g C_2H_6 to mols.

$$\frac{45.3 \text{ g } C_2H_6}{} \left| \frac{1 \text{ mol}}{\underset{\substack{\uparrow \\ 2 \text{ C} + 6 \text{ H}}}{30.069 \text{ g}}} \right. = \mathbf{1.51 \text{ mol } C_2H_6}$$

**H_2 = 2 H, O_2 = 2 O

#12 Notes III. Percent Composition

Ex. 1) Find the % composition of $C_6H_8O_6$ (ascorbic acid, which is Vitamin C)

6 C = 6 (12.011 g) = 72.066 g C
8 H = 8 (1.0080 g) = 8.064 g H
6 O = 6 (15.999 g) = 95.994 g O
176.124 g

$\% C = \frac{\text{mass of C}}{\text{molar mass}} \times 100 = \frac{72.066\text{ g}}{176.124\text{ g}} \times 100 = \mathbf{40.9\ \%\ C}$ Keep 3 or 4 digits!

$\% H = \frac{\text{mass of H}}{\text{molar mass}} \times 100 = \frac{8.064\text{ g}}{176.124\text{ g}} \times 100 = \mathbf{4.58\ \%\ H}$

$\% O = \frac{\text{mass of O}}{\text{molar mass}} \times 100 = \frac{95.994\text{ g}}{176.124\text{ g}} \times 100 = \underline{\mathbf{54.5\ \%\ O}}$

=99.98%

Ex. 2) Find the molar mass of a compound, if it is 23.9% oxygen. The compound contains 3 oxygen atoms in each molecule.

$\% O = \frac{\text{mass O}}{\text{molar mass}} \times 100$

$23.9\% = \frac{3\ (15.999\text{ g})}{(\text{mm})} \times 100$

23.9 (mm) = 4799.7
mm = 2.01×10^2 g/mol

IV. Empirical Formula

-is the simplest whole # ratio of atoms in a compound.

Molecular Formula (Real Formula)	Empirical Formula
N_2H_4	$N_1H_2 = NH_2$
$AlCl_3$	$AlCl_3$
$C_6H_{12}O_6$	CH_2O
$(NH_4)_2C_4H_4O_4 = N_2H_{12}C_4O_4 \rightarrow$	$NH_6C_2O_2$

Ex. 1a) Find the empirical formula of a compound containing 3.57 g Sc and 1.91 g O.

$Sc_\#O_\#$
↑ ↑
mols

i) Find mols: $\frac{3.57\text{ g Sc}}{} \Big| \frac{1\text{ mol}}{44.956\text{ g}} = 7.94109 \times 10^{-2}$ mol Sc

$$\frac{1.91\text{ g O}}{} \left| \frac{1\text{ mol}}{15.999\text{ g}} \right. = 1.193824\text{ X}10^{-1}\text{ mol O}$$

ii) Divide by the smallest: $\frac{7.94109\text{ X}10^{-2}\text{mol Sc}}{7.94109\text{ X}10^{-2}} = 1.00$

$$\frac{1.193824\text{ X}10^{-1}\text{ mol O}}{7.94109\text{ X}10^{-2}} = 1.50335 = 1.50$$

$Sc_{1.00}\ O_{1.50}$

iii) If necessary, multiply to make whole #'s: X2 X2 $\rightarrow$ $\mathbf{Sc_2O_3}$

Ex. 1b) What is the molecular formula, if the molar mass is 413.7 g/mol?

Sc_2O_3 = 2 Sc + 3 O = 137.909 g/mol

$\frac{\text{molar mass}}{\text{empirical mass}} = \frac{413.7\text{ g/mol}}{137.909\text{ g/mol}} = 3$ 3 times bigger, so Sc_2O_3 X3 = $\mathbf{Sc_6O_9}$

Ex. 2a) Find the empirical formula of a compound containing 37.7 % Na, 23.0 % Si and ? % O.

The percents must add up to 100%, so
100% – 37.7% Na – 23.0% Si = 39.3 % O

Assume we have a 100 g sample of the compound: 37.7 % of 100 g = 37.7 g Na
23.0 % of 100 g = 23.0 g Si and 39.3 % of 100 g = 39.3 g O

$$\frac{37.7\text{ g Na}}{} \left| \frac{1\text{ mol}}{22.990\text{ g}} \right. = 1.6398434\text{ mol Na} \quad / 8.189133\text{ X}10^{-1} = 2$$

$$\frac{23.0\text{ g Si}}{} \left| \frac{1\text{ mol}}{28.086\text{ g}} \right. = 8.189133\text{ X}10^{-1}\text{ mol Si} \quad / 8.189133\text{ X}10^{-1} = 1$$

$$\frac{39.3\text{ g O}}{} \left| \frac{1\text{ mol}}{15.999\text{ g}} \right. = 2.4564035\text{ mol O} \quad / 8.189133\text{ X}10^{-1} = 3$$

$\mathbf{Na_2SiO_3}$

Ex. 2b) What is the molecular formula, if the molar mass is 244 g?

Na_2SiO_3 = 122.062 g/mol $\frac{244\text{ g}}{122.062\text{ g/mol}} = 2$

Na_2SiO_3 X2 = $\mathbf{Na_4Si_2O_6}$

#13 Notes V. Hydrates

-water is incorporated inside the crystalline solid.

$FeSO_4 \cdot 7\ H_2O$ iron II sulfate heptahydrate

$Co(NO_3)_2 \cdot 6\ H_2O$ cobalt II nitrate hexahydrate

Ex. 1) Find the formula of the hydrate, if it contains 9.77 g $CuCl_2$ and 2.62 g H_2O.

$$\frac{9.77\ g\ CuCl_2}{} \Big| \frac{1\ mol}{134.452\ g} = 7.26653 \times 10^{-2}\ mol\ /\ 7.26653 \times 10^{-2} = 1$$

(Divide by the smallest, like empirical formula.)

$$\frac{2.62\ g\ H_2O}{} \Big| \frac{1\ mol}{18.015\ g} = 1.454343 \times 10^{-1}\ mol\ /\ 7.26653 \times 10^{-2} = 2$$

$CuCl_2 \cdot 2\ H_2O$

VI. Balancing Chemical Reactions

$H_2SO_{3(aq)} \rightarrow H_2O_{(l)} + SO_{2(g)}$

Reactants — Products

(react in the reaction) — (are produced/formed in the reaction)

aq = aqueous, s = solid, cr = crystalline solid, l = liquid, g = gas

Steps:

1) Put reactants on the left side of the arrow and the products on the right.
2) Balance the elements by changing the coefficients at the front of the compounds, until both sides are equivalent. (Do not change subscripts or put numbers into the compound!!)
   $H_2O \neq H_3O$ $H_2O \neq H_22O$

a) Balance metals first {(+) part of the compounds}.
b) Balance N or S.
c) Balance H or O.
d) Save for last whatever element is all over.

Ex. 1) $N_2O_5 \rightarrow NO_2 + O_2$

2 N	**2**(1N) = 2 N	Fix N, O is everywhere.
5 O	2 + 2 = 4 O	

$N_2O_5 \rightarrow \underline{\mathbf{2}}\ NO_2 + O_2$

2N	2N	
5 O	4 + 2 = 6 O	We need more O on the left, so try doubling the N_2O_5.

$\underline{\mathbf{2}}\ N_2O_5 \rightarrow 2\ NO_2 + O_2$

4 N	$\underline{\mathbf{2}}$ (2N) = 4 N	Refix N.
10 O	4 + 2 = 6 O	

$2\ N_2O_5 \rightarrow \underline{\mathbf{4}}\ NO_2 + O_2$

4N	4N	
10 O	8 + 2 = 10 O	balanced

Ex. 2) $Cr(NO_3)_3 + NaOH \rightarrow Cr(OH)_3 + NaNO_3$

1 Cr	1 Cr	
3 N	$\underline{\mathbf{3}}$(1 N) = 3	Fix N or H, not O (everywhere)
9 + 1 = 10 O	3 + 3 = 6 O	
1 Na	1 Na	
1 H	3 H	

$Cr(NO_3)_3 + NaOH \rightarrow Cr(OH)_3 + \underline{\mathbf{3}}\ NaNO_3$

1 Cr	1 Cr	
3 N	3 N	
9 + 1 = 10 O	3 + 9 = 12 O	
$\underline{\mathbf{3}}$(1Na) = 3	3 Na	Fix Na or H
1 H	3 H	

$\mathbf{Cr(NO_3)_3 + \underline{3}\ NaOH \rightarrow Cr(OH)_3 + 3\ NaNO_3}$

1 Cr	1 Cr	
3 N	3N	
9 + 3 = 12 O	3 + 9 = 12 O	
3 Na	3 Na	
3 H	3 H	balanced

Ex. 3) $O_{2(g)} + As_2S_{3(s)} \rightarrow As_4O_{6(s)} + SO_{2(g)}$

2 O	6 + 2 = 8 O	
$\underline{\mathbf{2}}$(2 As) = 4	4 As	Fix As or S.
3 S	1 S	

$O_{2(g)}$ + **2** $As_2S_{3(s)} \rightarrow As_4O_{6(s)} + SO_{2(g)}$

2 O	6 + 2 = 8 O
4As	4 As
6 S	**6**(1 S) = 6 S

$O_{2(g)} + 2\ As_2S_{3(s)} \rightarrow As_4O_{6(s)}$ + **6** $SO_{2(g)}$

9(2 O) = 18 O	6 + 12 = 18 O
4As	4 As
6 S	6 S

9 $\mathbf{O_{2(g)} + 2\ As_2S_{3(s)} \rightarrow As_4O_{6(s)} + 6\ SO_{2(g)}}$ balanced

** if odd/even problem, multiply everything by 2
** if no clue, try adding a 2 somewhere, then a 3, then a 4 (trial and error)

#14 Notes VII. 5 Types of Chemical Reactions

1) Decomposition (one compound falls apart to 2 or more compounds)

$Ca(OH)_2 \rightarrow CaO + H_2O$

2) Synthesis (2 or more compounds combine to form one compound)

$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$

3) Combustion (burning)

$\mathbf{Compound + O_2 \rightarrow CO_2 + H_2O}$

Combustion of C_3H_6:

$2\ C_3H_6 + 9\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$

4) Single Displacement (elements and compounds, one element replaces another)

$Cl_2 + 2\ KI \rightarrow 2\ KCl + I_2$

K moves over to the Cl, leaving I alone

5) Double Displacement (all compounds, 2 elements/groups replace each other)

$Al_2(SO_4)_3 + 3\ Ca(OH)_2 \rightarrow 2\ Al(OH)_3 + 3\ CaSO_4$

Al moves to OH

Ca moves to SO_4

VIII. Stoichiometry

Steps:

1) Write the balanced chemical reaction.
2) Write a conversion equation.
    a) Find the mols of the compound with known mass.
    b) Use the mol ratio (in the balanced reaction) between the 2 compounds you are interested in.
    c) Find the grams of the compound you are looking for.

The only time you look at the balanced reaction is for step 2b.!!

Ex. 1) How many grams of HCl will react with 44.7 g $Ca(OH)_2$?

$$2\,HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2\,H_2O$$

$$\frac{44.7\text{ g }Ca(OH)_2}{} \left|\frac{1\text{ mol }Ca(OH)_2}{74.094\text{ g }Ca(OH)_2}\right|\frac{2\text{ mol HCl}}{1\text{ mol }Ca(OH)_2}\left|\frac{36.461\text{ g HCl}}{1\text{ mol HCl}}\right. = \mathbf{44.0\text{ g HCl}}$$

Ex. 2) What would be the minimum amount of carbon monoxide used, if 80.3 g iron were produced?

$$Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$$

$$\frac{80.3\text{ g Fe}}{} \left|\frac{1\text{ mol Fe}}{55.847\text{ g Fe}}\right|\frac{3\text{ mol CO}}{2\text{ mol Fe}}\left|\frac{28.01015\text{ g CO}}{1\text{ mol CO}}\right. = \mathbf{60.4\text{ g CO}}$$

#15 Notes IX. <u>Limiting Reagent</u>

-is the reactant that makes the least amount of product.

How many cars?

3 steering wheels + 20 tires → 3 cars

↓ ↓

makes 3 cars makes 5 cars (4 tires per car)

(limiting reagent)

3 steering wheels + 8 tires → 2 cars

↓ ↓

makes 3 cars makes 2 cars

(limiting reagent) ****smallest amount will be the answer due to the limiting reagent.**

Ex. 1a) Given 76.5 g iron III oxide and 45.0 g carbon monoxide, find the mass of iron produced.

$$\underset{76.5\text{ g}}{\overset{\uparrow}{Fe_2O_3}} + \underset{45.0\text{ g}}{\overset{\uparrow}{3\,CO}} \rightarrow \underset{?\text{ g}}{\overset{\uparrow}{2\,Fe}} + 3\,CO_2$$

$$\frac{76.5\text{ g }Fe_2O_3}{} \left| \frac{1\text{ mol }Fe_2O_3}{159.691\text{ g }Fe_2O_3} \right| \frac{2\text{ mol Fe}}{1\text{ mol }Fe_2O_3} \left| \frac{55.847\text{ g Fe}}{1\text{ mol Fe}} \right. = 53.5\text{ g Fe}$$

$$\frac{45.0\text{ g CO}}{} \left| \frac{1\text{ mol CO}}{28.01015\text{ g CO}} \right| \frac{2\text{ mol Fe}}{3\text{ mol CO}} \left| \frac{55.847\text{ g Fe}}{1\text{ mol Fe}} \right. = 59.8\text{ g Fe}$$

53.5 g Fe (the answer will be the least.)

Ex. 1b) How much CO did not react?

$$\frac{53.5\text{ g Fe}}{} \left| \frac{1\text{ mol Fe}}{55.847\text{ g Fe}} \right| \frac{3\text{ mol CO}}{2\text{ mol Fe}} \left| \frac{28.01015\text{ g CO}}{1\text{ mol CO}} \right. = 40.3\text{ g CO } \textbf{did react}$$

45.0 g originally given

<u>- 40.3 g reacted</u>

4.7 g unreacted

*Use answer and go back to reactant that was not limiting to find how much of it actually reacts.

** Don't forget: D = m/v (if have density and volume, find mass)

#16 Notes X. Percent Yield

-shows the efficiency of a reaction.

The limiting reagent problems give the amount of product that **should** be produced (the **theoretical** yield).

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Ex. 1) Looking back at Asst #15 Ex.1 : An amount of Fe_2O_3 and CO was given.
A limiting reagent problem determined the amount of Fe produced (53.5 g).
What is the % yield, if only 47.4 g Fe was produced in the experiment?

$$\% \text{ yield} = \frac{47.4 \text{ g}}{53.5 \text{ g}} \times 100 = 88.6\ \%$$

XI. Stoichiometry and Empirical Formula

Ex. 1) Glucose contains C, H, and O. Combustion of 11.9 mg glucose produces 17.5 mg CO_2 and 7.15 mg H_2O. The molar mass of glucose is 180 g/mol. Calculate the empirical and molecular formula.

$$\underset{\substack{\uparrow \\ 11.9 \text{ mg} \\ (0.0119 \text{ g})}}{\text{C, H, O}} + O_2 \rightarrow \underset{\substack{\uparrow \\ 17.5 \text{ mg} \\ (0.0175 \text{ g})}}{CO_2} + \underset{\substack{\uparrow \\ 7.15 \text{ mg} \\ (0.00715 \text{ g})}}{H_2O} \qquad \frac{1 \times 10^{-3} \text{ g}}{1 \text{ mg}} =$$

$$\frac{0.0175 \text{ g } CO_2}{} \left| \frac{1 \text{ mol } CO_2}{44.00915 \text{ g } CO_2} \right| \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} \left| \frac{12.011 \text{ g C}}{1 \text{ mol C}} \right. = \mathbf{0.00478\ g\ C}$$

$$\frac{0.00715 \text{ g } H_2O}{} \left| \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \right| \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \left| \frac{1.0080 \text{ g H}}{1 \text{ mol H}} \right. = \mathbf{8.00 \times 10^{-4}\ g\ H}$$

0.0119 g C,H,O - 0.00478 g C - 0.000800 g H = 0.00632 g O

$$\frac{0.00478 \text{ g C}}{} \left| \frac{1 \text{ mol C}}{12.011 \text{ g C}} \right. = 3.98 \times 10^{-4} \text{ mol C} / 3.95 \times 10^{-4} = 1$$

$$\frac{8.00 \times 10^{-4} \text{ g H}}{} \left| \frac{1 \text{ mol H}}{1.0080 \text{ g H}} \right. = 7.94 \times 10^{-4} \text{ mol H} / 3.95 \times 10^{-4} = 2$$

$$\frac{0.00632 \text{ g O}}{} \left| \frac{1 \text{ mol O}}{15.999 \text{ g O}} \right. = 3.95 \times 10^{-4} \text{ mol O} / 3.95 \times 10^{-4} = 1$$

$C_1H_2O_1$

180 / 30.026 g/mol = 6 **$C_6H_{12}O_6$**

All gases are diatomic: N_2, H_2, O_2 except Noble Gases: He, Ne etc.

End of Notes (Assignments #17-18 are Review Assignments. There are no notes for these assignments.)