**Chemistry II—Chapter 9 Notes**

**Stoichiometry**

- The Arithmetic Of Equations:

 - when you know the amount of one substance in a reaction, you can

 calculate the amounts of the other reactants consumed or products

 formed

- this method is called STOICHIOMETRY

- Greek stoichion (amount) and metry (to measure)

 - Interpreting Chemical Equations:

 N2(g) + 3 H2(g)  🡪 2 NH3(g)

 2 atoms N 6 atoms H = 2 atoms N / 6 atoms H

 1 molecule N2 3 molecules H2 = 2 molecules NH3

 10 molecules N2 30 molecules H2 = 20 molecules NH3

 1(6.02 x 1023 molecules) 3(6.02 x 1023 molecules) = 2(6.02 x 1023 molecules)

 1 mole N2 3 moles H2 = 2 moles NH3

 28 g N2 3 x 2g = 6 g H2 = 2 x 17g = 34 g NH3

 34 g reactants = 34 g products

- Chemical Calculations:

 - in ALL stoichiometric calculations you MUST begin with the

 BALANCED CHEMICAL EQUATION!!!

- Mole-Mole Calculations:

 - if you know the balanced equation, and the number of moles

 of any substance in a reaction, you can find the number of

 moles of any other substance in the reaction

- need a MOLE RATIO—a conversion factor relating moles of

 one substance to moles of another substance in a balanced

 equation

- use the COEFFICENTS from the balanced equation to get the

 mole ratio

- use the FACTOR LABEL METHOD to solve problems

N2 + 3 H2 🡪 2 NH3

*How many moles of NH3 are formed if 2.7 moles of N2 react?*

$$\left(\frac{2.7 mol N\_{2}}{1}\right)\left(\frac{2 mol NH\_{3}}{1 mol N\_{2}}\right)=5.4 mol NH\_{3}$$

 - Mole-Mass calculations:

 - need to use a mole ratio and then convert moles to grams

 - Al(s) + O2(g) 🡪 Al2O3(s)

 *How many grams of Al2O3 will form if 4.5 moles of Al react?*

4 Al(s) + 3 O2(g) 🡪 2 Al2O3(s)

 4.5 mol ? g

$$\left(\frac{4.5 mol Al}{1}\right)\left(\frac{2 mol Al\_{2}O\_{3}}{4 mol Al}\right)\left(\frac{102 g Al\_{2}O\_{3}}{1 mol Al\_{2}O\_{3}}\right)$$

$$=230 g Al\_{2}O\_{3}$$

 - Mass-Mole calculations:

 - need to convert grams to moles then use the mole ratio

 - Fe(s) + O2(g) 🡪 Fe2O3(s)

 *How many moles of Fe are needed to produce 62 grams of Fe2O3?*

4 Fe(s) + 3 O2(g) 🡪 2 Fe2O3(s)

 ?? mol 62 g

$$\left(\frac{62 g Fe\_{2}O\_{3}}{1}\right)\left(\frac{1 mol Fe\_{2}O\_{3}}{159.8 g Fe\_{2}O\_{3}}\right)\left(\frac{4 mol Fe}{2 mol Fe\_{2}O\_{3}}\right)$$

 $=0.78 mol Fe$

 - Mass-Mass calculations:

 - need to convert mass to moles, then use mole ratio, then

 convert moles to mass

- H2(g) + O2(g) 🡪 H2O(l)

*How many grams of H2O will form if 10. g H2 react?*

2 H2(g) + O2(g) 🡪 2 H2O(l)

 10.g ? g

$$\left(\frac{10. g H\_{2}}{1}\right)\left(\frac{1 mol H\_{2}}{2.0 g H\_{2}}\right)\left(\frac{2 mol H\_{2}O}{2 mol H\_{2}}\right)\left(\frac{18.0 g H\_{2}O}{1 mol H\_{2}O}\right)$$

$$=90 g H\_{2}O$$

**Mass A**

**(grams)**

 **Moles A**

 **(mol)**

 **Moles B**

 **(mol)**

 **Mass B**

 **(grams)**

- Limiting Reagent and Percent Yield:

 - not all reactions complete so 100% of all reactants are used up to

 form products

- THEORETICAL YIELD—the calculated amount of a product that

 will form if the reaction goes to 100% completion (which rarely

 happens!)

- any chemical reaction will continue until ALL of one reactant is

 used up

- whenever ALL of one reactant is gone, the reaction stops and you

 may have other reactants “leftover”

- LIMITING REAGENT—the reactant that limits the amount of

 product that can be formed (the first one to be used up in the

 reaction)

- EXCESS REACTANT—the reactant that is leftover when the

 limiting reagent is used up

- Mg(s) + HCl(aq) 🡪 MgCl2(aq) + H2(g)

*If 5.00 g Mg reacts with 6.00 g HCl,*

 *a) what is the limiting reagent?*

 *b) what is the excess reagent?*

 *c) what is the theoretical yield of H*2 *in grams?*

 *d) how much of the excess reactant is left?*

* Mg(s) + 2 HCl(aq) 🡪 MgCl2(aq) + H2(g)

 5.00 g 6.00g

$$\left(\frac{5.00 g Mg}{1}\right)\left(\frac{1 molMg}{24.3 g Mg}\right)\left(\frac{1 mol H\_{2}}{1 mol Mg}\right)\left(\frac{2.02 g H\_{2}}{1 mol H\_{2}}\right)$$

$$=0.416 g H\_{2}$$

$$\left(\frac{6.00 g HCl}{1}\right)\left(\frac{1 molHCl}{36.5 g HCl}\right)\left(\frac{1 mol H\_{2}}{2 mol HCl}\right)\left(\frac{2.02 g H\_{2}}{1 mol H\_{2}}\right)$$

$$=0.167 g H\_{2}$$

* So 5.00 g Mg 🡪 0.416 g H2
* And 6.00 g Mg 🡪 0.167 g H2

**So we ONLY make 0.167 g H2 !! Because after we make 0.167 g of H2 there is NO MORE HCl so HCl is the limiting reactant and Mg is the excess reactant**

$$\left(\frac{6.00 g HCl}{1}\right)\left(\frac{1 molHCl}{36.5 g HCl}\right)\left(\frac{1 mol Mg}{2 mol HCl}\right)\left(\frac{24.3 g Mg}{1 mol Mg}\right)$$

$$=2.00g Mg USED$$

**So 5.00 g – 2.00 g = 3.00 g Mg LEFT!!**

 - Percent Yield—a comparison of the ACTUAL yield (what an

 experimenter actually gets from a reaction) and the

 THEORETICAL yield (the calculated amount you should get if the

 reaction goes to 100% completion)

 ACTUAL

Percent Yield = ------------------------- x 100

 THEORETICAL

 - CaCO3(s) 🡪 CaO(s) + CO2(g)

*What is the theoretical yield of CaO if 24.8 grams of CaCO3 is heated?*

*What is the % yield if only 13.1 g CaO are produced?*

CaCO3(s) 🡪 CaO(s) + CO2(g)

 24.8 g ? g

$$\left(\frac{24.8 g CaCO\_{3}}{1}\right)\left(\frac{1 mol CaCO\_{3}}{100. g CaCO\_{3}}\right)\left(\frac{1 mol CaO}{1 mol CaCO\_{3}}\right)\left(\frac{56.0 g CaO}{1 mol CaO}\right)$$

$$=13.9 g CaO$$

 ***theoretical yield***

$$\%= \frac{actual }{theoretical} x 100$$

$$\%= \frac{13.1 g CaO}{13.9 g CaO} x 100=94.2\%$$