Chapter 9: Chemical Quantities DRAFT UPDATE

These Notes are to SUPPLIMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test!

The author is providing these notes as an addition to the students reading the text book and listening to the lecture. Although the author tries to keep errors to a minimum, the student is responsible for correcting any errors in these notes.

Chem Equations:

Start with a Balanced Equation

Coefficients give the relative numbers of molecules

CO gas	+	2 H_2 gas	→	CH_3 OH liq [Methanol]
1 molecule	+	2 molecules	→	1 molecule
6.022×10^{23}	+	$2 * 6.022 \times 10^{23}$	→	6.022×10^{23} molecules

Remember 6.022×10^{23} = 1 mole = Avagadro's Number

1 Mole	+	2 Moles	→	1 Mole

Mole to Mole relation

2 H_2O → 2 H_2 + O_2 decomposition of water

2 moles of water give 2 moles of H_2 and one mole of O_2

Determine what 4 moles of water will give [how much hydrogen and oxygen?]

Determine what 5.8 moles of water will give [how much hydrogen and oxygen?]

$$\frac{2 \text{ Moles } H_2O}{5.8 \text{ Moles } H_2O} = \frac{2 \text{ Moles } H_2}{\mathbf{X \text{ Moles } H_2}}$$

$$X = \frac{5.8\ H_2O * 2\ H_2}{2\ H_2O} = 5.8\ H_2O$$

More Questions: How do you get 2 moles of Oxygen etc, etc, etc

Mole ratios in calculations

Propane burns $C_3 H_8 + 5 O_2 \rightarrow 3 CO_2 + 4 H_2O$ Organics burn using Oxygen and generate CO_2 & H_2O

How many moles of Oxygen is needed for 4.3 moles of propane

1 mole of propane uses 5 moles of oxygen, setup a ratio:

$$\frac{1 \text{ Moles Propane}}{5 \text{ Moles Oxygen}} = \frac{4.3 \text{ Moles Propane}}{\mathbf{X \text{ Moles Oxygen}}}$$

$$\mathbf{X} \text{ Moles Oxygen} = \frac{4.3 \text{ Moles Propane} * 5 \text{ Moles Oxygen}}{1 \text{ Moles Propane}} = \mathbf{21.5 \text{ M Oxygen}}$$

Mass Calculations

Problem #1:

2 Al + 3 I_2 → 2 Al I_3 [Aluminum Iodide]

2 moles + 3 moles → 2 moles

Mw: Al = 26.98 g/mole I = 126.9 g/mole

35. g Al is reacted, how much Iodine is needed to use all of the Aluminum?

First Calculate how many moles of Al we have:

$$\frac{35.\text{ g Al}}{26.98\text{ g/mole Al}} = 1.30\text{ mole Al}$$

Since it takes 3 moles of I_2 [that's 6 atoms of Iodine] for each 2 moles of Al:

$$\text{Amount of Iodine} = 1.30\text{ mole} * \frac{3\text{ mole }I_2}{2\text{ mole Al}} = 1.95\text{ mole }I_2$$

Now Calculate how many grams of Iodine that 1.95 mole I_2

1.95 mile I_2 * 126.9 g / mole I_2 = 495. g Iodine

Or you could have figured the whole thins out the <u>RedNeck Way: [This is NOT in the Book]</u>
Take the number of grams given over the total molecular weight of that reactant set that equal to X over the total molecular weight of what your looking for:

Try to do these in an Organized manner. Put what is given on the top. What your suppose to find is a "?". Put the Balanced equation next and the molecular weight of each "Group" under that "Group"

35. g Al	**?**			**Put givens on the Top**
2 Al +	3 I_2	→	2 Al I_3	**Balanced Equation**
2 * 26.98 g/mole Al	6 * 126.9 g / mole			**Molecular Weights**

$$\frac{35.\text{ g Al}}{2 * 26.98\text{ g/mole Al}} = \frac{X}{3\text{ mole }I_2 * 2\text{ }I_2/I * 126.9\text{ g / mole }I_2}.$$

$$X = \frac{(35.\text{ g Al}) * (3\text{ mole }I_2 * 2\text{ }I_2/I * 126.9\text{ g / mole }I_2)}{2 * 26.98\text{ g Al /mole}} = 494\text{ g Iodine}$$

Now let's figure out how much Aluminum Iodide we get:

$$\frac{35.\text{ g Al}}{2 * 26.98\text{ g/mole Al}} = \frac{X}{815.4\text{ g/mole 2 Al }I_3}. \qquad X = 36.21 = 36.\text{ g Al }I_3$$

Next Example [Note this is a Redox Reaction, but we will not worry about that in Chem 1025 for burning of Organics to produce Carbon Dioxide and Water. You are responsible for Redox Reactions with other materials.

<u>Problem #2:</u>

Propane from you barbecue grill burns to from Water and Carbon Dioxide

$C_3H_8 + 0_2 \rightarrow CO_2 + H_2O$

But, it's not balanced:

$C_3H_8 + 5\ 0_2 \rightarrow 3\ CO_2 + 4\ H_2O$

Ok, now suppose we burn 96.1 g of Propane, how much Oxygen is required?
Let's do it the RedNeck Way. First Calculate the Mw of each:

C_3H_8				$5\ 0_2$	
C	3 * 12.01 =	36.03	g/mole	O	5 * 2 * 16.00 = **160.0 g/mole Ox**
H	8 *1.007 =	8.056	g/mole		
	Total	44.086	g/Mole = **44.09 g/mole Propane [Note the # of Significant Digits]**		

96.1	<u>**?**</u>			**Put givens on the Top**
C_3H_8 +	$5\ 0_2$ →	$3\ CO_2$ +	$4\ H_2O$	**Balanced Equation**
44.09 g/mole	160.0 g/mole			**Molecular Weights**

$$\frac{96.1 \text{ g Propane}}{44.09 \text{ g/mole Propane}} = \frac{X \text{ g Oxygen}}{160.0 \text{ g/mole Oxygen}}$$

X g Oxygen = 96.1 g Propane * 160.0 g/mole Oxygen / 44.09 g/mole Propane = 348.74 =

349 g Oxygen

Solve now for the amount of Carbon Dioxide and Water Generated?

Re-solve now for your barbecue grill. If 15 lbs of Propane is burned, now much Oxygen is required, how much Carbon Dioxide and Water is formed? Solve this in pounds [lbs] and in grams!

Stoichiometry is using a chemical equation to calculate the relative masses of reactants and products involved in a reaction.

Problem #3:

Lithium Hydroxide absorbs carbon dioxide. How much CO_2 can 1.00 x 10^3 g [or one kilogram] of LiOH absorb?

$2\,LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$

Li = 1 * 6.941 g/mole =	6.941 g	C = 1 * 12.01 g/mole =	12.01 g
O = 1 * 16.00 g/mole =	16.00 g	O = 2 * 16.00 g/mole =	32.00 g
H = 1 * 1.008 g/mole =	1.008 g	Mw =	44.01 g / mole of CO2
Mw =	23.949		
Mw =	23.95 g / mole LiOH		

Calculate the number of moles of LiOH used:

$$\frac{1.00 \times 10^3 \text{ g}}{23.95 \text{ g / Mole}} = 41.8 \text{ moles LiOH}$$

Since the mole ration is 2 moles of LiOH to 1 mole of CO_2, then the # of moles of CO_2 is:

$$41.8 \text{ mole LiOH} * \frac{1 \text{ Mole } CO_2}{2 \text{ mole Li OH}} = 0.9 \text{ mole } CO_2$$

The number of grams of CO_2 is the number of moles * the Mw of CO_2:

$0.9 \text{ mole } CO_2 * 44.01\text{g} = 920. \text{ g } CO_2$

Or the RedNeck way of doing it:

1.00 x 10^3 g		?				**Put givens on the Top**
2 LiOH	+	CO_2 ->	Li_2CO_3	+	H_2O	**Balanced Equation**
2 * 23.95 g/mole		44.01 g / mole				**Molecular Weights**

$$\frac{1.00 \times 10^3 \text{ g LiOH}}{2 \text{ mole LiOH} * 23.95 \text{ g/mole LiOH}} = \frac{X \text{ g CO2}}{44.01 \text{ g / mole of } CO_2}$$

$$X = \frac{1.00 \times 10^3 \text{ g} \quad \text{LiOH} * 44.01 \text{ g / mole of } CO_2}{2 \text{ mole LiOH} * 23.95 \text{ g/mole LiOH}} = 918.789 = 919 \text{ g } CO_2$$

You now solve this for the amount Lithium Carbonate and Water we can get from this reaction.

Self Check 9.4: Hydrofluoric Acid is used to etch glass by the following unbalanced equation:

$HF + SIO_2 \rightarrow SiF_4 + H_2O$

How much [what mass] of HF is needed to etch 5.68 g of glass [silica]?

How much water is produced?

Extra Question: Baking soda can be taken to calm an acid stomach [in the olden days]. If one table spoon [5.00 g] of baking soda is used, how much carbon dioxide is generated?

$NaHCO_3 + HCl \rightarrow NaCl + H_2CO_3 \rightarrow H_2O + CO_2$

Note: if one mole of CO_2 occupies 22.4 liters, how much CO_2 volume is generated?

Example 9.6 Baking Soda $NaHCO_3$ and Milk of Magnesia [$Mg(OH)_2$] are often used to calm an upset stomach by the following unbalanced equations:

$NaHCO_3 + HCl \rightarrow NaCl + H_2CO_3 \rightarrow H_2O + CO_2$

$Mg(OH)_2 + HCl \rightarrow MgCl_2 + H_2O$

If one gram of both Baking Soda and Milk of Magnesia are used, which will neutralize the most stomach acid?

<u>Limiting Ractant:</u> Reactants are not mixed in Stoichiometric Quantities, you must determine which is the limiting reactant and which is in excess.

Example of a sandwhich. You have 5 slices of bread and 2 slices of cheese, how many sandwhiches can you make using 2 slices of bread and 1 slice of cheese per sandwhich?

5 slices of bread / 2 slices of bread per sandwhich = 2.5 sandwhich

2 slices of cheese / 1 slice of cheese per sandwhich = 2 sandwhich

Take the smallest amount, you can make 2 sandwhiches

Example of a car. You have 4 cars chassis, 12 tires, how many finished cars can you make?

4 cars chassis / 1 chassis per car = 4 cars

12 tires / 4 tires per car = 3 cars

Take the smallest amount, you can make 3 cars.

Problem #4: **Burn Methane:** $CH_4 + H_2O \rightarrow 3H_2 + CO$

You have 249 g methane. How much water is required to react with all of that methane?

CH_4			**H_2O**		
1 C	1 * 12.00 g/mole =	12.00 g/mole	2 H	2 * 1.008 g/mole =	2.016 g/mole
4 H	4 * 1.008 g/mole =	4.032 g/mole	1 O	1 * 16.00 g/mole =	16.00 g/mole
		16.032 g/mole			18.016 g/mole
		16.03 g/mole			**18.02 g/mole**

249 g	<u>**?**</u>				
CH_4 +	H_2O	->	$3H_2$	+	CO
16.03 g/mole	18.02 g/mole				

$$\frac{249\ g}{\mathbf{16.03\ g/mole}} = \frac{X\ g}{\mathbf{18.02\ g/mole}}$$

What if you have 300. g HOH?

First you determine which compound is present in the smallest molar amount

$\frac{249.\ g\ CH_4}{16.03\ g/mole} = 15.533 =$ **15.5 moles CH_4** $\frac{300\ g\ H_2O}{18.02\ g/mole} = 16.648 =$ **16.6 moles H_2O**

The **CH_4** is in the smaller amount, so we use this for all of our calculations. So now let's calculate how much of each of the components is required or produced.

249 g	<u>**?**</u>		<u>**?**</u>		<u>**?**</u>
CH_4 +	H_2O	->	$3H_2$	+	CO
16.03 g/mole	18.02 g/mole		6.048 g/mole		28.00 g/mole

$\frac{249\ g\ CH_4}{16.03\ g/mole\ CH_4} = \frac{X}{6.048\ g/mole\ 3H_2}$ X = 93.946 = 93.9 g Hydrogen

$\frac{249\ g\ CH_4}{16.03\ g/mole\ CH_4} = \frac{X}{28.00\ g/mole\ CO}$ X = 434.93 = 435 g CO

$\frac{249\ g\ CH_4}{16.03\ g/mole\ CH_4} = \frac{X}{18.02\ g/mole\ H_2O}$ X = 279.91 = 280 g Water

But – we supplied 300 g Water [look above]. So we have 300 g – 280 g = 20 g excess of water!

All of the methane is consumed before the water runs out? WE JUST PROVED IT!

Problem #5:

Example 9.7 How much ammonia is formed from reacting 25.0 kg of nitrogen gas with 5.00 kg of hydrogen gas: $N_2 + H_2 \rightarrow NH_3$

Ok, this is a Redox Reaction, so: $\overset{0}{N_2} + \overset{0}{H_2} \rightarrow \overset{+3\ +3}{NH_3}$

$N_2 \quad - \quad -6e^- \rightarrow 2 N^{+3}$

$H_2 \rightarrow 2H^+ + 2e^-$

We must multiply the bottom half reaction by 3 so we gain and loose the same amount, 6 electrons

$N_2 \quad - \quad -6e^- \rightarrow 2 N^{+3}$

$3 H_2 \rightarrow 6H^+ + 6e^-$

$N_2 + 3 H_2 \rightarrow 2\ NH_3$

Now we can do the math and find out which is the limiting reagent:

25.0 kg	5.00 kg		**Put givens on the Top**
N_2 +	$3 H_2$ →	$2\ NH_3$	**Balanced Equation**
28.02 g/mole	6.048 g/mole		**Molecular Weights**

$\frac{25.0 \text{ kg } N_2}{28.02 \text{ g/mole}}$ = 892.21 = **892 moles** $\qquad \frac{5.00 \text{ kg } 3 H_2}{6.048 \text{ g/mole}}$ = 826.72 = **827 moles**

The smallest number is the 827 moles of Hydrogen, so this is the limiting reagent and the one that we do all of your calculations by.

?	5.00 kg	**?**	**Put givens on the Top**
N_2 +	$3 H_2$ →	$2\ NH_3$	**Balanced Equation**
28.02 g/mole	6.048 g/mole	34.07 g/mole	**Molecular Weights**

$\frac{5.00 \text{ kg } 3 H_2}{6.048 \text{ g/mole } H_2} = \frac{X}{34.07 \text{ g/mole } 2\ NH_3}$ $\qquad$ X = 28.166 = 28.2 kg Ammonia or NH_3

$\frac{5.00 \text{ kg } 3 H_2}{6.048 \text{ g/mole } H_2} = \frac{X}{28.02 \text{ g/mole } N_2}$ $\qquad$ X = 23.164 = 23.2 kg of Nitrogen or N_2

Now this reaction used up 23.2 kg of Nitrogen. But we started with 25.0 kg of Nitrogen. So we have left 25.0 kg – 23.2 kg = 1.8 kg of Nitrogen.

Rules for Solving a Limiting Reagent Problem:

1. Write and Balance the Equation

2. Put the Given's on the top of each compound

3. Put the Complete Molecular Weigh under each compound

4. Convert the masses given to moles. Moles = Grams / Complete Molecular Weight

5. Use number of moles in the reactants equation, determine which is the limiting reagent.

The smallest number is the Limiting Reagent.

6. Determine the amount of each of the products that is formed.

7. Determine how much of the Limiting Reagent is used.

8. Determine how much of an excess there is of the Limiting Reagent.

% Yield

Theoretical Yield - amount of product calculated to get – or the 100% yield

Actual yield is called the Percent Yield is what you actually get in the lab.

When you determine the amount of product formed from a reaction, that is the 100% yield. All of the amounts of products calculated above are the Theoretical Yield or 100% yield.

Most reactions do not give 100% yield, they give LESS! Multiply this percentage by the amount of yield you have calculated and that is amount of yield.

Problem #1 above:

If we only obtained 30. G of AlI3 our Percentage Yield is: $\frac{30.\ g}{36,\ g}$ * 100% = 83 % Yield

Problem #4 above:

If we only obtained 350. g of CO, our Percentage Yield is: $\frac{350.\ g}{435,\ g}$ * 100% = 80.5 % Yield

Problem #4 above:

If this reaction is known to generate a 55% yield of Hydrogen gas, how much Hydrogen do we get?
0.55 * 93.9 g Hydrogen = 51.6 g of Hydrogen

Example 9.9 68.5 kg of Carbon Monoxide is reacted with 8.60 kg of hydrogen to produce 3.57×10^4 g of Methanol by the following equation:

$H_2 + CO \rightarrow CH_3\text{-}OH$

A. What is the theoretical yield of Methanol?
B. What is the actual percentage yield of this reaction?