Name: ______________________________

Unit 3

Advanced Topics in Bonding, Moles & Stoichiometry

- Apply the mole concept to substances.
- Determine the number of particles and the amount of substance (in moles).
- Define the terms relative atomic mass (A_r) and relative molecular mass (M_r).
- Calculate the mass of one mole of a species from its formula.
- Solve problems involving the relationship between the amount of substance in moles, mass and molar mass.
- Deduce chemical equations when all reactants and products are given.
- Identify the mole ratio of any two species in a chemical equation.
- Calculate theoretical yields from chemical equations.
- Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given.
- Solve problems involving theoretical, experimental and percentage yield.

Determining Empirical Formula & Molecular Formula from Percent Composition

The molecular formula for a compound can be determined from the percentage composition by assuming the sample has a mass of 100 g. Using the percentages, the number of grams out of 100 can be determined for each component. This can be converted to moles by dividing by the Gram Formula Mass (GFM). The mole ratio and empirical formula can be determined by dividend each number of moles by the smallest number of moles. The atomic masses are added together to find the empirical formula mass. The empirical formula mass is divided into the molecular weight to find the number of times 'n' the formula is repeated. Finally, 'n' is multiplied by the empirical formula to find the molecular formula.

Sample Problem

Find the molecular formula for a compound composed of 5.9% hydrogen and 94.1% oxygen and having a molecular weight of 34 amu.

Step 1: Assume a 100 g sample

Step 2: Find the mass of each element in the sample

mass of H = 5.9 % of 100 g = 5.9 g
mass of O = 94.1 % of 100 g = 94.1 g

Step 3: Convert grams to moles

moles of H $= \frac{5.9g}{1g/mol}$ = 5.9 moles

moles of O $= \frac{94.1g}{16g/mol}$ = 5.9 moles

Step 4: Find the mole ratio by dividing both numbers by the smaller number

5.9 ÷ 5.9 = 1 H
5.9 ÷ 5.9 = 1 O
empirical formula = HO

Step 5: Find the empirical formula mass

atomic mass of H = 1
atomic mass of O = 16
EFM = 17

Step 6: Find the number of times, "n," the empirical formula is repeated and multiply through

$\frac{M.W.}{EFM}$ = n = $\frac{34}{17}$ = 2

molecular formula $(HO)_n$ = $(HO)_2$ = H_2O_2

Use the Percent Composition equation from Table T of your reference tables.

1. What is the percent composition of *each element* in the following compound?

   a) MgOH

2. Calculate the percentage of water in the following hydrate (the • means there are x many water molecules for every 1 of the corresponding compound.

   a) $BeBr_2 \bullet 6H_2O$

3. A strip of tin (Sn) weighing 8.1 grams is heated in a stream of oxygen (O_2) until it is converted to an oxide. The mass of the oxide is 9.2 g.

   a) What is the percentage composition of the resultant oxide?

   b) Is the formula of this compound SnO or SnO_2? (show your work)

4. Find the empirical formula of the following compound (use the % composition given).

   a) Na = 43.4%, C = 11.3%, O = 45.3%

5. Find the molecular formula of the following compound.

   a) C = 93.75%, H = 6.25%; Molecular Mass = 128 g/mol

Practice with Word Equations

- *Write the word descriptions below as chemical equations, making sure each chemical formula is correct. Remember diatomic molecules and that compounds must have a neutral overall charge.*
- *Balance the equation with the smallest whole-number coefficients.*

1. Scandium metal and mercury (II) nitrate react to form scandium nitrate and liquid mercury.

2. Gallium bromide and fluorine gas react to form gallium fluoride and bromine gas.

3. Potassium phosphate and strontium chloride react to form strontium phosphate and potassium chloride.

4. Aluminum metal and hydrochloric acid react to form aluminum chloride and hydrogen gas.

5. Calcium hydroxide and nitric acid react to form calcium nitrate and water.

Gram → Gram Calculations
Mass/Mass Problems

With a balanced equation, a *Periodic Table*, and some knowledge of chemistry, you can figure out how much of any product will form from a given amount of reactant. Consider the following problem:

How much oxygen is needed to produce 27.0 g of water by burning hydrogen?

Step 1: Write a balanced equation
$2H_2 + O_2 \rightarrow 2H_2O$

Step 2: Calculate the GFM of the known and unknown.

O_2: $O = 16 \times 2 = 32$

H_2O: $H = 1 \times 2 = 2$; $O = 16 \times 1 = 16$; total 18

Step 3: Apply the factor label method

$$27\,g_{H_2O} \times \frac{1\,mol_{H_2O}}{18\,g_{H_2O}} \times \frac{1\,mol_{O_2}}{2\,mol_{H_2O}} \times \frac{32\,g_{O_2}}{1\,mol_{O_2}} = 24\,g_{O_2}$$

You will notice that, in using dimensional analysis (factor label method), you are first converting grams of the known material to moles, then moles of the known material to moles of the unknown material using a proportion from the coefficients of the balanced equation, and, finally, moles of the unknown material to grams.

Calculate the amount of material asked for in each of the following. A balanced equation is provided.

1. How many grams of oxygen will be produced from the decomposition of 155 grams of $KClO_3$?

   $2KClO_3 \rightarrow 2KCl + 3O_2$

2. How many grams of Zn will be needed to completely react with 45 g of HCl?

   $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

3. How many grams of hydrogen will be needed to react with 120 g of nitrogen according to the following?

   $N_2 + 3H_2 \rightarrow 2NH_3$

Molar Volume Calculations (Gases Only)

1. Calcium carbonate decomposes at high temperatures to form carbon dioxide and calcium oxide:

$$CaCO_{3(s)} \rightarrow CO_{2(g)} + CaO_{(s)}$$

How many grams of calcium carbonate will I need to form 2.58 liters of carbon dioxide?

2. Ethylene burns in oxygen to form carbon dioxide and water vapor:

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)}$$

How many liters of carbon dioxide can be formed if 2.25 liters of ethylene are consumed in this reaction?

3. Nitrogen monoxide and oxygen gas combine to form the brown gas nitrogen dioxide. How many mL of nitrogen dioxide are produced when 4.5 mL of oxygen reacts with an excess of nitrogen monoxide? Assume conditions of STP. (Remember 1L = 1000mL)

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

of Particle Calculations - Avogadro's Number

1. How many **molecules** of oxygen are produced by the decomposition of 6.54 g of $KClO_3$?

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

How many **atoms** of oxygen is this?

2. How many **molecules** of nitrogen monoxide are produced if 92.0 g of nitrogen dioxide is combine with excess water according to the following reaction? (Not balanced)

$$\_\_NO_{2(g)} + \_\_H_2O_{(l)} \rightarrow \_\_\ HNO_{3\ (g)} + \_\_NO_{\ (g)}$$

3. What is the mass of 99.0 x 10^{23} molecules of H_2?

4. How many **atoms** are contained in 10.7 g of Nitrogen gas?

Mixed Mole/Stoichiometry Problems

1. $N_2 + 3H_2 \rightarrow 2NH_3$

What volume of NH_3 at STP is produced if 25.0 g of N_2 is reacted with an excess of H_2?

2. $2KClO_3 \rightarrow 2KCl + 3O_2$

If 5.0 g of $KClO_3$ is decomposed, what volume of O_2 is produced at STP?

3. How many grams of KCl are produced in Problem 2?

4. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

What volume of hydrogen at STP is produced when 2.5 g of zinc react with an excess of hydrochloric acid?

5. $H_2SO_4 + 2NaOH \rightarrow 2H_2O + Na_2SO_4$

How many molecules of water are produced if 2.0 g of sodium sulfate are produced in the above reaction?

Limiting Reactants (Reagents)

You drop a piece of zinc into a beaker of hydrochloric acid. It begins to bubble furiously, but eventually it stops. you drop another piece of zinc into the acid. The bubbling begins anew, but again it stops. This time, you add more hydrochloric acid. Nothing happens. Obviously, each time the reaction stopped it was because you ran out of zinc. There was always plenty of hydrochloric acid. In fact, there as an excess of the acid. Zinc, on the other hand, was a limiting reactant. The balanced equation predicts the amounts of reactants needed to completely consume each other (stoichiometric quantities). If any of the reactants is in excess, the other(s) is (are) limiting reactants.

In stoichiometry problems where the amount of only one of the reactants is given, it is assumed either that there are stoichiometric amounts of all the reactants and products, or, at the very least, that all of the reactant for which the amount is specified is consumed because is is the limiting reactant. for problems in which the amount of more than one reactant is specified, you need to consider that there may be a limiting reactant. If there is a limiting reactant, you need to know which one it is and use it for your calculations, because excess, unconsumed reactants do NOT produce any product.

Using your notes from class, complete the following limiting reactant problem.

1. How much copper will precipitate when 10.0 g of granular zinc are added to a solution containing 16.0 g of aqueous copper II sulfate? What is in excess, and by how much? Determine which reactant is limiting during your calculations

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Percent Yield

In chemical reactions, the actual yield is usually less than the theoretical yield due to side reactions and other complications. The theoretical yield is the amount of product formed when the limiting reactant is completely consumed. It is the maximum amount of product that can be produced. The actual yield is often expressed as a percentage of the theoretical yield called the percent yield.

In these problems, determine the limiting reagent and thus the theoretical yield and then calculate the percent yield by dividing the actual yield by the theoretical yield.

1. 325 g of iron III oxide are treated in a blast furnace with 150 g of carbon monoxide to form 130 g of pure iron. What is the percent yield?

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

2. 585 g of nitrogen are reacted with 145 g of hydrogen under conditions of high temperature and pressure in the presence of a catalyst to produce 105 g of ammonia. What is the percent yield?

$$N_2 + 3H_2 \rightarrow 2NH_3$$