Honors Chemistry
Chapter 9 Stoichiometry pgs. 299-318.

Chapter 9 Section 1: Introduction to Stoichiometry pgs. 299-303.

Objectives:
1. **Describe the importance of the mole ratio in stoichiometric calculations.**
2. **Write a mole ratio relating two substances in a chemical equation.**

Vocabulary: Define the following.
1. **composition stoichiometry--**

2. **reaction stoichiometry--**

3. **mole ratio--**

There are two types of mass relationships used in stoichiometry. The first ratio called the composition stoichiometry deals with the mass of the elements within the compounds of the equation and the second ratio called the reaction stoichiometry compares the mass of the reactants and products in relationship to each other.

Stoichiometry

Stoichiometry is the study of the quantitative (measurable) relationships that exist in chemical formulas and chemical reactions. In other words, stoichiometry utilizes all of the knowledge gained in the previous chapters and adds to this mix the concept of how ratios affect the outcome. The two ratios encountered in chemical problems are the ratios found in a (1) chemical formula, example $MgCl_2$ has a ratio of 1Mg : 2Cl, and the (2) molar ratios found in the chemical equation as represented by the coefficients in the equation.

For instance look at the following chemical reaction.

$$FeS + 2HCl \rightarrow H_2S + FeCl_2$$

Each compound has a ratio (number 1 above); FeS shows a ratio of 1Fe : 1S, H_2S has a ratio of 2H : 1S, etc.

The equation shows other ratios based on the mole (number 2 above). 1 mol of FeS reacts with 2 mol HCl to produce 1 mol of H_2S and 1 mol of $FeCl_2$.

As you may have guessed, the coefficients are the number of moles.

Note: If the chemical formula of any substance is written incorrectly or if the equation is incorrectly balanced, then you cannot proceed.

Correct chemical formulas and balancing is your prior knowledge and the first part of stoichiometry.

There are several types of stoichiometric problems. In each of the problems the substances must be either in moles or changed to moles. ****The essential ingredient is therefore the number of moles of each substance. ****All of these problems use dimensional analysis to obtain moles or grams.

<u>Problem Types:</u>

1. **mole to mole.**
2. **mole to mole to grams.**
3. **grams to moles to moles.**
4. **grams to moles to moles to grams.**

<u>Mole Ratio</u>

The mole ratio is the amount of moles found in the balanced chemical equation. For example in the chemical equation $4NH_3(g) + 6NO(g) \rightarrow 5N_2(g) + 6H_2O(l)$,
the mole ratio of NO to N_2 would be written as: $\frac{6 \text{ mol NO}}{5 \text{ mol } N_2}$
and the mole ratio of H_2O to NO would be written as: $\frac{6\ H_2O}{6\ NO}$

******The writing of these ratios as conversion factors shows the relationship of the moles of one compound to moles of another compound in the equation. These conversion factors written as fractions allow one to flip the conversion factor over if needed to cancel like terms. The conversion factors show for instance that 6 mols of NO are needed to produce 5 moles of N_2 and likewise 6 moles of H_2O are produced for every 6 moles of NO as reactants.**

Please note the following.

The number of significant figures in the answer is determined by the number of significant figures of <u>any measured quantity</u> in the problem. Any molar masses obtained from the periodic table are considered as infinite and need <u>not</u> be considered in determining significant figures in the answer.

<u>Molar Mass</u>

The molar mass as discussed in a previous chapter is the mass in grams of an element per 1 mole of that substance as read from the periodic table and for a compound it is the sum of all the individual elements of that element. In the previous chemical equation, the molar masses of the 4 compounds involved are:

NH_3 = 14.01g/mol (N) + 3 x 1.01g/mol (H) = 17.04g/mol NH_3
NO = 14.01g/mol (N) + 16.00g/mol (O) = 30.01g/mol NO
N_2 = 14.01g/mol x 2 = 28.02g/mol N_2
H_2O = 1.01g/mol(H) x 2 + (O) 16.00g/mol = 18.02 g/mol H_2O

Remember that in dimensional analysis conversion factors can be flipped so that the appropriate label cancels.

From the chemical equation given on the previous page, how many grams of NH_3 are there in 4 moles of NH_3?

Solution: $\frac{4 \text{ mol } NH_3}{1} \times \frac{17.04 \text{ grams } NH_3}{1 \text{ mol } NH_3} = 68.16 \text{ grams } NH_3$

Try the following.

1. **How many grams are in 6 moles of NO?**

2. **How many grams are in 5 moles of N_2?**

3. How many grams are in 6 moles of H_2O?

Chapter 9 Section 2 Ideal Stoichiometric Equations pgs. 304-311.

Objectives:

1. **Calculate the amount in moles of a reactant or product from the amount in moles of a different reactant or product**
2. **Calculate the mass of a reactant or product from the amount in moles of a different product.**
3. **Calculate the amount in moles of a reactant or product from the mass of a different reactant or product**
4. **Calculate the mass of a reactant or product from the mass of a different reactant or product.**

Ideal stoichiometric calculations assume that all reactants are completely changed into the products predicted by the chemical equation. Not all chemical reactions occur in an ideal manner. In many instances not all the reactants are completely converted into the predicted products. ****All stoichiometric calculations begin with a correctly balanced chemical equation. Theoretical calculations predict the maximum amount of product that can be produced.

Conversions of Quantities in Moles

The following is an example of a mole to mole stoichiometric problem.

Iron (II) sulfide reacts with hydrochloric acid to produce hydrogen sulfide and iron (II) chloride. How many moles of iron (II) chloride are produced if 0.32 mol of iron (II) sulfide is present?

Step 1: Write the skeleton equation.

$$FeS + HCl \rightarrow H_2S + FeCl_2$$

Step 2: Balance the equation.

$$FeS + 2HCl \rightarrow H_2S + FeCl_2$$

Step 3: Solve the problem using dimensional analysis.

$$\frac{0.32 \text{ mol FeS}}{1} \times \frac{1 \text{ mol } FeCl_2}{1 \text{ mol FeS}} = 0.32 \text{ mol } FeCl_2$$

What if the above question instead asked: How many moles of iron (II) chloride are produced from 1.38 moles of hydrochloric acid?

Answer:

$$\frac{1.38 \text{ mol HCl}}{1} \times \frac{1 \text{ mol } FeCl_2}{2 \text{ mol HCl}} = 0.69 \text{ mol } FeCl_2$$

Always show your work!!!!

Answers without work are considered incomplete and receive no credit!!!

Try the following:

1. Lead reacts with hydrochloric acid to produce lead (II) chloride and hydrogen gas. How many moles of HCl are needed to completely react with 0.36 mol of lead?

2. Lithium reacts with oxygen to form lithium oxide. How many moles of lithium oxide will be formed if 5 moles of lithium react?

3. Hydrogen peroxide (H_2O_2) decomposes to form water and oxygen gas. How many moles of O_2 will result if 5 moles of hydrogen peroxide decompose?

<u>Conversions of Amounts in Moles to Mass</u>

Example: Iron(II) sulfide reacts with hydrochloric acid to produce hydrogen sulfide and iron(II) chloride.

How many moles of iron(II) chloride are produced if 25.3 grams of hydrochloric acid reacts with the iron(II) chloride?

The skeleton is written first (step 1) and then balanced (step 2).

Step 1: $FeS + HCl \rightarrow H_2S + FeCl_2$

Step 2: $FeS + 2HCl \rightarrow H_2S + FeCl_2$

Step 3: Solve the problem by adding the given mass and using dimensional analysis.

$$\frac{25.3\text{g HCl}}{1} \times \frac{1 \text{ mol HCl}}{36.46\text{g HCl}} \times \frac{1 \text{ mol } FeCl_2}{2 \text{ mol HCl}} = 0.0395\text{mol } FeCl_2$$

Try the following.

1. Lead(II) nitrate reacts with sodium iodide to produce sodium nitrate and lead(II) iodide. How many moles of lead(II) iodide are produced from 152.7 grams of lead(II) nitrate?

2. Barium chloride reacts with sodium phosphate to produce barium phosphate and sodium chloride. How many moles of sodium chloride are produced from 46.8 grams of sodium phosphate?

Mass to Mass Calculations

Example: What mass of lead(II) iodide will be produced when 16.4 grams of lead(II) nitrate is reacts with potassium iodide to produce lead(II) iodide and potassium nitrate?

$$Pb(NO_3)_2 \quad + \quad 2KI \quad \rightarrow \quad PbI_2 \quad + \quad 2\ KNO_3$$

Solution:

$$\frac{16.4\ g\ Pb(NO_3)_2}{1} \times \frac{1\ mol\ Pb(NO_3)_2}{331g\ Pb(NO_3)_2} \times \frac{1\ mol\ PbI}{1\ mol\ Pb(NO_3)_2} \times \frac{461g\ PbI_2}{1\ mol\ PbI_2} = 22.8g\ PbI$$

Try the following.

1. **When sodium chloride reacts with silver nitrate, silver chloride and sodium nitrate are produced. What mass of AgCl is produced from 75.0g of $AgNO_3$?**

2. **Aluminum reacts with iron(II) oxide to produce iron and aluminum oxide. What mass of aluminum oxide is produced when 2.3 grams of aluminum reacts with iron(II) oxide?**

3. **Lithium nitride reacts with water to produce ammonia and lithium hydroxide. Determine the mass of lithium hydroxide produced when 0.38 grams of lithium nitride reacts.**

4. **Aluminum bromide reacts with chlorine gas to produce bromine and aluminum chloride. How many grams of aluminum chloride will be produced from 92 grams of Cl_2?**

Section 3: Limiting Reactants and Percent Yield pgs 312-318

Objectives:

1. **Describe a method for determining which of two reactants is a limiting reactant.**
2. **Calculate the amount in moles or mass in grams of a product, given the amounts in moles or masses in grams of two reactants, one of which is in excess.**

3. Distinguish between theoretical yield, actual yield, and percentage yield.
4. Calculate percentage yield, given the actual yield and quantity of a reactant.

<u>Vocabulary</u>: Define the following.

1. limiting reactant--

2. excess reactant--

3. theoretical yield--

4. actual yield--

5. percentage yield--

Sometimes because of the stoichiometric equation there is not enough or too much of one substance in a chemical reaction. The amount of the products formed by a reaction is dependent on the amount of reactants available.

These types of situations present the student with two new chemistry terms.

<u>Stoichiometric proportions</u> - The quantities of reactants are available in the exact ratio found in the equation.

<u>Nonstoichiometric proportions</u> - There is more of one reactant than can be used. It is these type of reactions in which the limiting reactant must be found.

<u>Identifying Limiting Reactants</u>

The reactant that limits the amount of product formed in a chemical reaction is the <u>limiting reactant</u>. The other reactant or reactants will be found in <u>excess</u> (left over).

***********The quantities of products formed in a reaction are always determined by the quantity of the limiting reactant.***

In most problems involving limiting reactants, the masses of two substances are given. It cannot be known from the amounts given for each substance in the problem which substance limits the reaction because of the molar ratios present in the problem.

In general, a limiting reactant problem consists of 2 separate mass-mass problems. Use the following steps to solve.

Step 1: Write the balanced equation.
Step 2: Observe the product or products formed. If there are more than one product, pick the one with the easiest molar mass.
Step 3: Enter the given mass in dimensional analysis to obtain the mass of the product determined in the problem.
Step 4: Repeat step 3 with the second given mass for the same product.

Step 5: The reactant that produces the smallest chosen product is the limiting reactant.

Example: Identify the limiting reactant when 2.20 g of magnesium reacts with 4.50 L of oxygen at STP to produce magnesium oxide.

Step 1: $2\,Mg + O_2 \rightarrow 2\,MgO$

Step 2: There is only one product MgO.

Step 3: $\frac{2.20g\ Mg}{1} \times \frac{1\ mol\ Mg}{24.3g\ Mg} \times \frac{2\ mol\ MgO}{2\ mol\ Mg} \times \frac{40.3g\ MgO}{1\ mol\ MgO} = 3.65g\ MgO$

Step 4: $\frac{4.50L\ O_2}{1} \times \frac{1\ mol\ O_2}{22.4L\ O_2} \times \frac{2\ mol\ MgO}{1\ mol\ O_2} \times \frac{40.3g\ MgO}{1\ mol\ MgO} = 16.2g\ MgO$

Step 5: Mg produces 3.65g MgO and 4.50L O_2 produces 16.2g MgO. The smallest quantity of MgO produced is 3.65g which was produced from the 2.20g of Mg. Therefore, Mg is the limiting reactant.

1. **Lead(II) nitrate reacts with potassium iodide to produce lead(II) iodide and potassium nitrate according to the equation:** $Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$
   What is the limiting reactant when 16.40 grams of lead(II) nitrate is added to 28.50 grams of potassium iodide?

2. **Copper combines with silver nitrate to form copper(II) nitrate and silver by the reaction shown as** $Cu + 2\,AgNO_3 \rightarrow 2\,Ag + Cu(NO_3)_2$
   What is the limiting reactant using a reaction in which 3.50 gram of copper is added to a solution of containing 6.00 grams of silver(I) nitrate?

Percent Yield

In chemical reactions the amounts of product predicted by the stoichiometric calculation and the actual product obtained in the lab may not always be the same. The amount predicted by the equation is the expected yield, and the amount obtained in the laboratory is the actual yield. The formula for obtaining the percent yield is:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Expected Yield}} \times 100\%$$

Example: A particular reaction is expected to produce 2.6L of oxygen gas. In reality, the reaction only produces 1.9L of oxygen gas. What is the percent yield?

$$\text{Percent Yield} = \frac{1.9L}{2.6L} \times 100\% = 73\%$$

Many times the percent yield is combined with a limiting reactant question, such as the following.
Problem: Determine the percent yield for the reaction between 6.92g of K and 4.28 g of O_2 if 7.36g of K_2O is produced.

Solution:

$$4\,K + O_2 \rightarrow 2\,K_2O$$

$$\frac{6.92\text{ g K}}{1} \times \frac{1\text{ mol K}}{39.1\text{ g K}} \times \frac{2\text{ mol }K_2O}{4\text{ mol K}} \times \frac{94.2\text{ g }K_2O}{1\text{ mol }K_2O} = 8.34\text{ g }K_2O$$

$$\frac{4.28\text{ g }O_2}{1} \times \frac{1\text{ mol }O_2}{32.0\text{ g }O_2} \times \frac{2\text{ mol }K_2O}{1\text{ mol }O_2} \times \frac{94.2\text{ mol }K_2O}{1\text{ mol }K_2O} = 25.2\text{ g }K_2O$$

The limiting reactant is the K. This means that the expected yield is 8.34 g K_2O.

$$\text{Percent Yield} = \frac{7.36\text{ g}}{8.34\text{ g}} \times 100\% = 88.2\%$$

1. **Problem: A piece of copper with a mass of 5.00 grams is placed in a solution of silver(I) nitrate containing excess $AgNO_3$. If 15.2 grams of silver are produced, what is the percent yield for this reaction?**
   The balanced equation is $Cu + 2\,AgNO_3 \rightarrow 2\,Ag + Cu(NO_3)_2$

2. **Problem: When octane (C_8H_{18}) is burned in oxygen, carbon dioxide and water are produced. If 320.0 grams of octane are burned and 392.0 grams of water are recovered, what is the percent yield of the experiment? The equation is $2\,C_8H_{18}\;25\,O_2 \rightarrow 16\,CO + 18\,H_2O$**