Chapter 9 – Stoichiometry
Part 1 – Notes: Introduction to Stoichiometry

Objectives: Calculate the number of moles of any other reactant or product from the number of moles of one or more of the reactants or products in a chemical reaction.
Calculate the amount of reactant required or the product formed in various processes (chemical or nonchemical).
Interpret balanced chemical equations in terms of moles, representative particles, masses, and volumes at STP.
Determine the mole ratio used for various calculations and explain why and when to use such a mole ratio.
Identify, define, and explain: stoichiometry, mole ratio, and conversion factor.
Text Reference: Section 9.1 and Section 9.2 (part) – pages 237-246

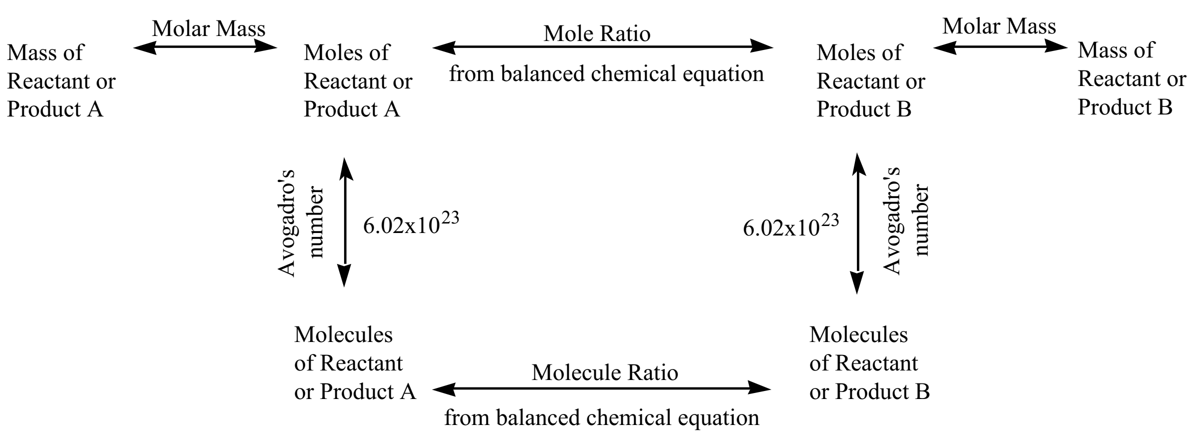
Interpreting Chemical Equations: Let's examine the equation behind the reaction of nitrogen with hydrogen to produce ammonia.

Symbolic equation	
Picture	
Atoms	
Molecules	
Moles	
Mass	
Conserved mass	
Volume (STP)	

Stoichiometry involves the calculation of quantities of any substances involved in a chemical reaction from the quantities of other substances

The balanced equation gives the ratios of formula units of all the reactants and products of a chemical reaction. It also gives the corresponding ratios of moles of reactants and products.

Examine the conversions that may be obtained from a balanced chemical equation:



There is, of course, no requirement that a chemist place exactly one mole of nitrogen and three moles of hydrogen in the reaction flask. *The equation supplies the* ***reacting ratio***. Ratios of coefficients from balanced chemical equations may be used as conversion factors for solving problems. This is a ***MOLE RATIO***.

Example 1: If you react sodium with chlorine, what is the balanced chemical equation?
When given a quantity of sodium and being asked about a quantity of chlorine, what ratio will you use?
When given a quantity of chlorine and being asked about a quantity of sodium chloride, what ratio will you use?

The ratio used to solve a specific problem will relate the substance that you are given and the substance that you want to find in the proper position so that the unit and the substance labels cancel out and you are left with the proper unit and label.

There are four basic steps involved in solving mole-mole problems:

Step 1: Write the balanced chemical equation for the given reaction.
Step 2: Show the problem specifications (the haves and needs).
Step 3: Show the mole proportion established by the equation that you will use.
Step 4: Solve the problem using unit analysis. Include all units and labels. You MUST use unit analysis.

Example 2: Calculate the number of moles of sodium atoms that will exactly react with 1.73 moles of chlorine molecules to from sodium chloride? The balanced equation is: $2\ Na + Cl_2 \rightarrow 2\ NaCl$

Essentially, all problems involving mole calculations are as simple as the ones in the above examples. A problem may *seem* more difficult if you are required to write and balance the equation before you perform the calculation. It is very important that the equation be written and balanced correctly; otherwise, the ratio or the compounds themselves may be wrong.

Example 3: Calculate the number of moles of aqueous sulfuric acid that must react completely with excess aqueous sodium hydroxide to produce 1.55 moles of sodium sulfate.

Remember the quantities involved in mole calculations are the quantities that ***react***, not necessarily the quantities that are *present*.

Example 4: A sample of 0.1500 moles of solid $KClO_3$ is heated gently over a period of time, and 0.0850 moles of the substance decomposes. Calculate the moles of oxygen gas produced.

Example 5: 65.00 g NaCl react with excess silver nitrate. What mass, in grams, of silver chloride is produced.

Chapter 9 – Stoichiometry

Part 1 – Assignment: Introduction to Stoichiometry

Solve the following problems. YOU MUST USE UNIT ANALYSIS AND SHOW ALL WORK, LABELS, ETC.

1. Balance this equation: C_2H_5OH + O_2 → CO_2 + H_2O

   Show that the balanced chemical equation obeys the law of conservation of mass.

2. Explain this statement: "Mass and atoms are conserved in every chemical reaction, but moles will not necessarily be conserved."

3. Carbon disulfide is an important industrial solvent. It is prepared by the reaction of coke (assume pure C) with sulfur dioxide. Carbon monoxide is the other product of this reaction.
   a. How many moles of carbon disulfide form when 2.70 moles of C reacts?

   b. How many moles of carbon are needed to react with 5.44 moles sulfur dioxide?

   c. How many moles of carbon monoxide form at the same time that 0.246 mole carbon disulfide form?

   d. How many moles of sulfur dioxide are required to make 118 mol carbon disulfide?

4. $4\ FeCr_2O_4 + 8\ K_2CO_3 + 7\ O_2 \rightarrow 2\ Fe_2O_3 + 8\ K_2CrO_4 + 8\ CO_2$
   a. How many grams of $FeCr_2O_4$ are required to produce 44.95 g CO_2?

   b. How many moles of oxygen are required to produce 100. g of Fe_2O_3?

   c. If 450. g $FeCr_2O_4$ react, how many moles of oxygen will be consumed?

   d. How many grams of Fe_2O_3 will be produced from 3.765 moles of $FeCr_2O_4$?

   e. How many grams of K_2CrO_4 are formed per gram of K_2CO_3 used? (This is exactly 1 g of K_2CO_3.)

Chapter 9 – Stoichiometry
Part 2 – Notes: Stoichiometry with Various Quantities

Objectives: Calculate how much of any reactant is involved in a chemical reaction given any quantity of other reactant or product, regardless of initial units used.
Construct mole ratios from balanced chemical equations and apply these ratios to mole-mole and other stoichiometric calculations, including mass, representative particles, and volumes of gases at STP.
Identify, define, and explain: mole ratio.

Text Reference: Section 9.2 – pages 242-250

Not only masses, but quantities in any units, may be used for stoichiometric calculations. The given quantities must be changed to moles. Just as a mass is a unit of measure of the number of moles of a reactant or product, the number of individual atoms, molecules, or ions involved in a chemical reaction may be converted to moles of reactant or product and used to solve a problem.

Also, the number of moles of individual atoms or ions of a given element within a compound may also be used to determine the number of moles of reactant or product. Even the density of a substance may be used to determine with mass that may then be turned into a quantity in moles.

Essentially, all problems involving mole calculations are as simple as the ones only involving moles. A problem may seem more difficult if you are required to write and balance the equation before you perform the calculation. It is very important that the equation be written and balanced correctly; otherwise, the ratio or the compounds themselves may be wrong. Make sure you realize that in order to go from substance A to substance B you need to use a ***MOLE RATIO*** *and in order to use a mole ratio, you need to have your quantity in moles. Keep this in mind and you will be fine.*

Example 1: How many oxygen ***molecules*** does it take to produce 70.0 mol of water through a synthesis reaction with hydrogen?

Example 2: Calculate the number of ***molecules*** of carbon dioxide produced by complete combustion of 25.0 g of gaseous C_2H_6?

Example 3: Calculate the number of ***moles*** of solid mercury (II) oxide that can be produced by the reaction of oxygen gas with 20.00 mL of liquid mercury (density = 13.6 g/mL).

Example 4: The quantities of nitrogen, phosphorus, and potassium in a fertilizer are critical to the fertilizer's function in helping crops grow. Calculate the number of individual nitrogen atoms in the ammonium phosphate produced by the reaction of excess aqueous ammonia (NH_3) with 13.7 moles of phosphoric acid?

Example 5: Using the above reaction and equation, how many (individual) ***hydrogen atoms*** (as part of the ammonium ion) are produced when 48.06 g of ammonia (NH_3) are reacted?

Example 6: Phosphorus (P_4) and hydrogen can be combined to form phosphine (PH_3). How many ***liters*** (at STP) of hydrogen are required to react with 125.0 g of phosphorus (P_4)?

Chapter 9 – Stoichiometry
Part 2 – Assignment: Stoichiometry with Various Quantities

Solve the following problems, show all work, formulas, equations, and YOU MUST USE UNIT ANALYSIS.

1. The hydrogen used for about 90% of the industrial synthesis of ammonia comes from the following reaction at high temperature: $CH_4\ (g) + H_2O\ (g) \rightarrow CO\ (g) + 3\ H_2\ (g)$
   Calculate the number of molecules of CH_4 required to produce 1.00 metric tom ($1.00\text{x}10^6$ grams) of H_2.

2. Calculate the volume of oxygen gas at STP produced when 25.0 g potassium chlorate decomposes.

3. You have $4.09\text{x}10^{22}$ atoms of chlorine in a sample of CCl_4. How many grams of $COCl_2$ may be prepared from your sample according to the following: $2\ CCl_4\ (l) + O_2\ (g) \rightarrow 2\ COCl_2\ (g) + 2\ Cl_2\ (g)$?

4. Calculate the number of chlorine atoms in the iron (III) chloride prepared by treating excess iron (II) chloride with 10.7 g of chlorine gas.

5. _____$K_2Cr_2O_7$ + _____KI + _____H_2SO_4 → _____$Cr_2(SO_4)_3$ + _____K_2SO_4 + _____I_2 + _____H_2O
   a. Calculate the number of moles of potassium dichromate needed to react with 2.56 moles potassium iodide.

b. Calculate the number of grams of chromium (III) sulfate that is produced when 65.85 g of water is also produced.

6. Tin (II) fluoride, formerly found in many toothpastes, is formed when solid tin reacts with gaseous hydrofluoric acid.

a. How many liters of HF are needed to produce 9.40 L of hydrogen gas at STP?

b. How many molecules of hydrogen are produced by the reaction of tin with 20.0 L of HF at STP?

c. How many grams of tin (II) fluoride can be made by reacting 7.42×10^{24} molecules of HF with tin?

7. Lithium nitride reacts with water to form ammonia (NH_3) and aqueous lithium hydroxide.

a. What mass of water is needed to react with 32.9 g lithium nitride?

b. Calculate the number of grams of lithium nitride that must be added to an excess of water to produce 15.0 L NH_3 at STP.

8. If the reaction below proceeds with 96.8% yield, how many kilograms of $CaSO_4$ are formed when 5.24 kg SO_2 reacts with excess $CaCO_3$ and O_2? Reaction: $2\,CaCO_3\,(s) + 2\,SO_2\,(g) + O_2\,(g) \rightarrow 2\,CaSO_4\,(s) + 2\,CO_2\,(g)$

9. Ethyl alcohol (C_2H_5OH) can be produced by the fermentation of glucose ($C_6H_{12}O_6$). If it takes 5.00 hours to produce 8.00 kg of alcohol, how many days will it take to consume 1.00×10^3 kg of glucose? (An enzyme is used.)

$$C_6H_{12}O_6 \rightarrow 2\,C_2H_5OH + 2\,CO_2$$

Chapter 9 – Stoichiometry
Part 3 – Notes: Limiting Reactant Problems

Objectives: Identify and use the limiting reagent in a reaction to calculate the maximum amount of product produced and the quantity of excess reactant.
Identify, define, and explain: limiting reagent and excess reagent.
Explain why one reactant would be used in excess.

Text Reference: Section 9.3 (part) 252-256

Until now, we have been dealing with ideal chemical reactions. Take the reaction:

$$2Mg + O_2 \rightarrow 2\ MgO$$

- From this reaction, we can see that 2 atoms of magnesium react with 1 molecule of oxygen to form 2 molecules of magnesium oxide.
- We can also see that 2 moles of magnesium react with 1 mole of oxygen to form 2 moles of magnesium oxide.
- We can also see that $1.20\text{x}10^{24}$ atoms magnesium react with $6.02\text{x}10^{23}$ molecules oxygen to form $1.20\text{x}10^{24}$ molecules MgO.

However, in reality, it is difficult to get every atoms of magnesium to come together with 1 molecule of oxygen so they will combine during the reaction. Consequently, it is common practice to add more than is necessary of one reactant to make certain all of the atoms or molecules of the other reactant will be involved in the chemical reaction.

EXAMPLE – Take the reaction of magnesium and oxygen above. The two reactants are being reacted in order to produce a specific amount of magnesium oxide. If the desired amount of magnesium oxide is two moles, we know that *ideally* two moles of magnesium should react with one mole of oxygen to produce the two moles of magnesium oxide; however, some atoms of magnesium may not combine with oxygen molecules and the reaction might only produce 1.95 mol MgO.

To increase the chance that all magnesium atoms undergo reaction and produce the desired 2.00 mol of MgO in the reaction, let's say we use four moles of oxygen gas with our original two moles of magnesium. Even though four moles of oxygen are used during the reaction, only one mole of oxygen will actually undergo reaction because only two moles of magnesium are present.

In our above example, the magnesium is the ***LIMITING REACTANT*** because the amount of magnesium used in this case *limits the amount of product that may be obtained from the reaction.*

Oxygen is the ***REACTANT IN EXCESS***. We know from the balanced equation that only one mole of oxygen will react for every two moles of magnesium. In this case there are three moles of oxygen molecules in excess.

Interpreting Chemical Reactions:

Symbolic equation

Microscopic Recipe

Macroscopic Recipe

Before Reaction

After Reaction

Excess Reactant: Limiting Reactant:

Product is detrermined by the . . .

Steps to solve a PROBLEM WITH A LIMITING REACTANT

1. Write the balanced chemical equation.
2. Write the problem specifications. (What do you have? What do you need?)
3. What is the molar mass of any substance that relates back to grams/
4. Convert all of the given substances to moles from their initial units. If they are in moles, leave them as they are.
5. Perform calculations and complete a *have/need* table for both the given reactants.
6. Identify the limiting reactant and the excess reactant.
7. Calculate the amount of product requested in the unit requested.
8. Calculate the excess reactant that remains.

Example 1: 176.92 g sodium are added to 825.65 g chlorine to produce sodium chloride. (A) Which reactant is limiting? (B) Which reactant is in excess? (C) How many grams of excess remains after the reaction is completed?

Example 2: 105.45 g calcium are reacted with 166.55 g copper (I) nitrate. (A) Determine which is the limiting reactant? (B) How many grams of calcium nitrate are produced? (C) How many atoms of copper are produced? (D) What is the mass of the excess reactant that remains?

Chapter 9 – Stoichiometry
Part 3 – Assignment: Limiting Reactant Problems

Complete the following. Show all work, set-ups, units, labels, etc. Be neat!!!

1. For each balanced chemical equation, identify the limiting and excess reactant for the given combination of reactants.

a.	$2\,Al + 3\,Cl_2 \rightarrow 2\,AlCl_3$	3.6 mol aluminum	5.3 mol chlorine
b.	$2\,H_2 + O_2 \rightarrow 2\,H_2O$	6.4 mol hydrogen	3.4 mol oxygen
c.	$2\,P_2O_5 + 6\,H_2O \rightarrow 4\,H_3PO_4$	0.48 mol P_2O_5	1.52 mol water
d.	$4\,P + 5\,O_2 \rightarrow P_2O_5$	14.5 mol phosphorus	18.0 mol P_2O_5

2. For each problem in question 1, determine the amount of product formed.

a. b.

c. d.

3. 1.21 moles zinc are added to 2.65 moles of hydrochloric acid. (A) Determine the limiting reactant? (B) What is the mass of the excess reactant that remains after the reaction is completed?

4. 10.45 g aluminum are reacted with 66.55 g copper (II) sulfate. (A) What is the limiting reactant? (B) What is the mass of the excess reactant that remains after the reaction is completed?

5. 12.00 moles sulfur dioxide combines with 256.00 g oxygen gas to form sulfur trioxide. (A) Which reactant is limiting? (B) How many moles of sulfur trioxide are produced? (C) How many molecules of sulfur trioxide are produced? (D) How many grams of excess reactant remain after the reaction is completed?

6. Textbook – page 263 – question 48. Record your answers here.

Chapter 9 – Stoichiometry
Part 4 – Notes: Theoretical and Percent Yield

Objective: Identify, define, and explain: theoretical yield, actual yield, percent yield.
Express the quantity of product obtained from a reaction as a percentage of what the reaction is theoretically capable of producing.
Calculate theoretical yield, actual yield, or percent yield, given appropriate starting information.
Explain why a reaction's actual yield is not equal to its theoretical yield.

Text Reference: Section 9.3 (part) – pages 256-259

When a quantity of a product is calculated from a quantity or quantities of reactants, as you have done in previous stoichiometric calculations, that quantity is called the **theoretical yield**. However, *you do not live in a theoretical world.*

When a reaction is run, *less product than the calculated amount is often obtained.*

Some of the products may stay in the solution in which the reaction was run, some side reactions may use up some of the reactants, the reaction may be stopped before it is completed, or the molecules may not collide with the right energy or at the right orientation to allow the reaction to occur.

No matter why, the fact is that ***reactions produce less product than the calculated theoretical amount – the actual amount produced is less than the theoretical amount.*** *No reaction may produce more than the theoretical yield.*

The ***PERCENT YIELD*** is defined as 100 times the ratio of the actual yield to the theoretical yield.

PERCENT YIELD = (ACTUAL YIELD / THEORETICAL YIELD) X 100

Example 1: Calculate the percent yield of a reaction if calculations indicate that 4.57 g of that product could be obtained but only 4.46 g of the product is actually obtained.

Example 2: Calculate the percent yield if 5.05 g of solid PCl_5 is obtained in a certain experiment in which 3.50 g of liquid PCl_3 is treated with excess gaseous Cl_2.

Example 3: Calculate the theoretical yield and the percent yield for an experiment if 4.30 g of liquid $SOCl_2$ is obtained from the reaction of 26.0 g of gaseous SO_2 and excess chlorine gas.

Question: What does the percent yield of a reaction measure?

Chapter 9 – Stoichiometry
Part 4 – Assignment: Theoretical and Percent Yield

Solve the following problems, showing all work, labeling, units, set-ups, etc. Be neat!

1. When 45.80 g of potassium carbonate are completely reacted with excess hydrochloric acid, 46.3 g of potassium chloride are produced. Carbon dioxide and water are also formed.
   a. Write a balanced chemical equation.

   b. Calculate the theoretical yield and percent yield of potassium chloride.

2. When 50.0 g silicon dioxide are heated with excess carbon, 32.2 g silicon carbide is produced.
   Reaction: $SiO_2 + 3\,C \rightarrow SiC + 2\,CO$
   a. What is the percent yield of this reaction?

   b. How many grams of CO are produced?

3. 3.01×10^{24} molecules of C_8H_{18} react with 1400.00 grams of oxygen in a combustion that produces 1100.00 g carbon dioxide. Calculate the theoretical yield of carbon dioxide.

4. A 1004.0-g sample of $CaCO_3$ that is 95.0% pure gives 225 L of carbon dioxide gas at STP when reacted with excess HCl. What is the density of carbon dioxide in g/L?

5. Textbook – page 263 – question 50. Record the work and answers here.