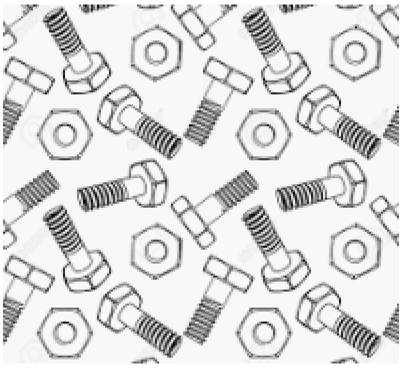
Subject Code Chem 1 **Chemistry 1**
Module Code 2.0 **Nomenclature of Inorganic Compounds**
Lesson Code 2.5 **Stoichiometry Part II (Limiting Reactant)**
Time Frame **30 minutes**

Components	Tasks	TA[1] (min)	ATA[2] (min)
Target 	After completing this module, you are expected to: **1.** determine the limiting and excess reactants. **2.** compute the theoretical and percent yield of the reaction. **3.** solve problems related to the concept of limiting reactant.	0.5	
Hook 	**How many nut and bolt pairs can be made from the picture below?**  Figure 1. Nuts and bolts (www. pinterest.com)	2.5	
Ignite 	**The Concept of Limiting Reactant** You can answer the question in the hook section by counting the number of bolts as well as the number of nuts. Because we only have eight nuts, the maximum number of bolt and nut pairs is eight. The number of nuts limits the number of pair that can be made. You can not have more than eight pairs since you only have eight nuts. If you cleary understood the answer to the nut and bolt question above, it would be very easy for you to gain mastery in the concepts discussed in this module. When a chemist carries out a reaction, the reactants are usually not present in exact stoichiometric amounts. Reactions are carried out with limited amount of one reactant and plentiful amounts of the others. Limiting reactant (or limiting reagent) *is the reactant that is entirely consumed when a reaction goes to completion*. The reactant that is NOT completely consumed is often referred as an *excess*	10	

[1]Time allocation suggested by the teacher.
[2]Actual time allocation spent by the student (for information purposes only).

reactant. Once the amount of limiting reactant is used up, the reaction stops.

Consider the reaction $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$, every mole of oxygen reacts with two moles of nitrogen monoxide. Therefore, 10 mol O_2 can react with 20 mol NO. However, suppose we had one tank containing **12mol O_2** and another with only **17mol NO**. If we let the two gases react, which reactant will be totally consumed? NO or O_2?

In order to find out the correct answer, let us have some calculations.

Assuming all O_2 is consumed,

$$12 \text{ mol } \cancel{O_2} \times \frac{2 \text{ mol NO}}{1 \text{ mol } \cancel{O_2}} = \mathbf{24 mol\ NO}.$$

Actual amount   mole ratio (theoretical, from equation)

24 mol NO are needed for all 12 mol O_2 to be consumed. But how many NO are available in the tank?

Only 17 molNO. So, it means that not all of our 12 mol O_2 will be consumed.
Assuming all NO is consumed,

$$17 \text{ mol } \cancel{NO} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } \cancel{NO}} = \mathbf{8.5 mol\ O_2}.$$

Only 8.5 mol O_2 are needed for all 17 mol NO to be consumed. But how many O_2 are available in the tank? It is 12 mol O_2. Therefore, O_2 is the *excess reactant*. It follows that NO is our *limiting reactant*.

How many mol of O_2 is in excess?

12 mol O_2 – 8.5 mol O_2 = **3.5 mol O_2**(unreacted O_2)

How many molof NO_2 then will be produced from 12 mol O_2 and 17 mol NO?

Solution: This where is the concept of limiting reactant is applicable.

$$17 \text{ mol } \cancel{NO} \times \frac{2 \text{ mol} NO_2}{2 \text{ } \cancel{\text{mol NO}}} = \mathbf{17 mol\ NO_2}$$

The moles of product are always determined by the starting moles of limiting reactant.

Therefore, the reactant that yields the least amount of product is the limiting reactant.

In summary:

	$2NO$	+	$O_2 \rightarrow$	$2NO_2$
Initial amount:	17 mol		12 mol	0 mol
Change(reaction):	-17mol		-8.5 mol	+17mol
Final amount:	0 mol		3.5 mol	17 mol

Using mass as the initial amount

Magnesium nitride, Mg_3N_2, is a product when magnesium metal meets nitrogen gas. (a) Calculate the amount of magnesium nitride that can be made in the reaction of 37.00 g of magnesium and 18.00 g of nitrogen? (b) How many grams of the excess reactant remain after the reaction? (page 102, Hill et. al,2005)

Solution:

Set-up the equation first: $Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$

Balance the equation: $3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$

Calculate: the mol of Mg_3N_2 using both Mg and N_2 as limiting reactants separately,

$$37.00\ g\ \cancel{Mg} \times \frac{1\ mol\ \cancel{Mg}}{24.305\ g\ \cancel{Mg}} \times \frac{1\ mol\ Mg_3N_2}{3\ mol\ \cancel{Mg}}$$

$= \mathbf{0.5074} mol\ Mg_3N_2$

$$18.00\ g\ \cancel{N_2} \times \frac{1\ \cancel{mol\ N_2}}{28.013\ \cancel{g\ N_2}} \times \frac{1\ mol\ Mg_3N_2}{1\ mol\ \cancel{N_2}}$$

$= 0.6426\ mol\ Mg_3N_2$

Since Mg yields the smaller amount of product,

Mg is the limiting reactant.

(a) $$0.5074\ mol\ \cancel{Mg_3N_2} \times \frac{100.93\ g\ Mg_3N_2}{1\ mol\ \cancel{Mg_3N_2}}$$

$= \mathbf{51.21\ g\ Mg_3N_2}$ (amount of product)

(b) $$0.5074\ mol\ \cancel{Mg_3N_2} \times \frac{1\ mol\ N_2}{1\ \cancel{mol\ Mg_3N_2}} \times \frac{28.013\ g\ N_2}{1\ \cancel{mol}\ N_2}$$

$= 14.21\ g\ N_2$ (consumed)

$18.00\ g\ N_{2(initial)} - 14.21\ g\ N_{2(consumed)} = \mathbf{3.79 g\ N_{2(excess)}}$

The Concept of Percent Yield

The amount of a product formed when the limiting reactant is completely consumed is called the **theoretical yield** of the product. This is the *maximum amount* that can be produced from the quantities of reactants used. The amount of product *actually* obtained, called the **actual yield**, is almost always less than (and can never be greater than) the theoretical yield. There are many reasons for this difference. One reason, it is not always possible to recover all of the product from the reaction mixture. The **percent yield** of a reaction relates actual and theoretical yield:

$$\text{Percent yield} = \frac{actual\ yield}{theoretical\ yield} \times 100\%$$

If the actual yield of Mg_3N_2 in the previous example is 47.97 g, what is the percent yield?

Percent yield = $\frac{47.97\text{g}}{51.21\text{g}} \times 100\ \% =$ **93.67 %**

Consider another example of the concept of percent yield:

If the actual yield of magnesium oxide is 2.05 g when 1.50 g of magnesium metal was burned in air, calculate the percent yield. The reaction is $2Mg + O_2 \rightarrow 2Mg$.
First we have to calculate the theoretical yield,

$$1.50\ \text{g}\ \cancel{Mg} \times \frac{1\ \cancel{mol\ Mg}}{24.31\ \cancel{g\ Mg}} \times \frac{2\ \cancel{mol}\ MgO}{2\ \cancel{mol\ Mg}} \times \frac{40.31\ g\ MgO}{1\ \cancel{mol}\ MgO} = \mathbf{2.49\ g}\ \text{MgO}$$

$$\%\ \text{yield} = \frac{2.05\ \cancel{g\ MgO}}{2.49\ \cancel{g\ MgO}} \times 100\% = \mathbf{82.3\ \%}$$

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Solve the following problems:

1. Calculate the number of moles of nitrogen monoxide gas that can be produced in the reaction of 3.50 mol ammonia gas and 4.90 mol oxygen gas ?
   $4NH_{3(g)} + 5O_{2(g)} \xrightarrow{\Delta} 4NO_{(g)} + 6H_2O_{(l)}$
   Answer: 3.50 moles of NO

2. There are many metals react with oxygen gas to form the metal oxide. For example, calcium reacts as follows:
   $2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$
   Calcium oxide that can be prepared from 4.40 g of Ca and 3.00 g of O_2.
   (a) What amount (mol) of calcium oxide can be produced from the given mass of Ca?
   (b) What amount (mol) of calcium oxide can be produced from the given mass of O_2?

(c) Which is the limiting reactant?

(d) How many grams of calcium oxide can be produced?

Answer: (a) 0.110 molCaO; (b) 0.188 molCaO; (c) calcium, (d) 6.16 g CaO

3. The depletion of ozone (O_3) in the stratosphere has been a matter of great concern among scientists in recent years. It is believed that ozone can react with nitric oxide (NO) that is discharged from the high-altitude jet plane, the SST. The reaction is

$$O_3 + NO \rightarrow NO_2 + O_2$$

If 0.735 g of ozone reacts with 0.687 g of nitrogen monoxide, how many grams of nitrogen dioxide will be produced? Which compound is the limiting reagent? Calculate the number of moles of the excess reagent remaining at the end of the reaction.

Answer: 0.704 g NO_2 ; O_3 ; 7.6 x 10^{-3}mol NO

4. Ammonia, NH_3, can be generated by heating together the solids ammonium chloride, NH_4C, land calcium hydroxide $Ca(OH)_2$. Calcium chloride, $CaCl_2$ and water are also formed. (a) If a mixture containing 44.0 g each of $NH_4Cl$and $Ca(OH)_2$ is heated, how many grams of NH_3will form? (b) Which reactant remains in excess, and in what mass?

Answer: (a) 14.0 g NH_3 ; (b) 13.5 g excess $Ca(OH)_2$

5. In the reaction of 288 g CCl_4 (calcium tetrachloride) with an excess of HF (hydrogen fluoride), 198 g CCl_2F_2(dichlorodifluoromethane) is obtained. What are the (a) theoretical, (b) actual, and (c) percent yields of this reaction?

$$CCl_4 + 2HF \rightarrow CCl_2F_2 + 2HCl$$

Answer: (a) 226 g CCl_2F_2 ; (b) 198 g CCl_2F_2 ; (c) 87.6 % yield

PART II GRADED ASSESSMENT (Must be done outside the 30 minute time limit of the module)

1. (a) When 1.50 mol of aluminum metal and 3.00 mol chlorine gas combine in the reaction , $2\ Al_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ AlCl_{3(s)}$, which is the limiting reactant? (b) How many moles of aluminium chloride are formed? (c) How many moles of the excess reactant remain at the end of the reaction?

2. A side reaction in the manufacture of rayon from wood pulp is $3CS_2 + 6NaOH \rightarrow 2Na_2CS_3 + Na_2CO_3 + 3H_2O$ How many grams of Na_2CS_3 are produced in the reaction of 93.2 mL of liquid CS_2 (d = 1.26g/mL) and 2.66 mol NaOH?

3. Consider the reaction

$$MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$$

If 0.92 mole of manganese dioxide and 47.4 g of hydrochloric acid react, which reagent will be used up first? How many grams of Cl_2 will be produced?

	4. Calculate the maximum numbers of moles and grams of iodic acid (HIO_3) that can form when 624 g of iodine trichloride (ICl_3) reacts with 120.5 g of water: $ICl_3 + H_2O \rightarrow ICl + HIO_3 + HCl$ (unbalanced) 5. When heated, lithium reacts with nitrogen to form lithium nitride: $6Li_{(s)} + N_{2(g)} \rightarrow 2Li_3N_{(s)}$ What is the theoretical yield of lithium nitride in grams when 11.8 g of lithium metal are heated with 34.5 g nitrogen gas? If the actual yield of lithium nitride is 8.88 g, what is the percent yield of the reaction?		
Knot 	The limiting reactant is the reactant that is completely consumed in a reaction. It is present in the smallest stoichiometric amount. Its quantity determines the theoretical quantity of the products formed. Other reactants are said to be present in excess. In some stoichiometry problems the limiting reactant must first be identified through a preliminary calculation. The amount of product obtained in a reaction (the actual yield) may be less than the maximum possible amount (the theoretical yield). The ratio of the two multiplied by 100 percent is expressed is expressed as the percent yield.	2	

References

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Chang, R. and Overby, J. (2011).*General Chemistry: The Essential Concepts*, 6th Ed. New York: McGraw-Hill

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