Session 6: LECTURE OUTLINE (SECTION M2 pp F88 - F90)

I. The Limiting Reagent
   a. Stoichiometric amounts
   b. Non-Stoichiometric amounts
   c. Limiting reagent (L.R)

II. How Do We Determine Which Is The L.R?
   a. Available Ratio
   b. Required ratio
   c. Example

III. Percent Yield
   a. Theoretical yield
   b. Actual yield
   c. Examples

Suggested problems: p F90 M.2A, M.2B

p F93

M.4, M.6, M.11

THE LIMITING REAGENT CONCEPT

Consider the reaction illustrated by the following chemical equation:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

The coefficients of the species in a balanced chemical equation show the ratio in which substances react and are produced. In every reaction, we can have one of two situations:

- One in which STOICHIOMETRIC amounts of reagents react. This means that the mole ratio of reactants is EXACTLY AS SHOWN BY THE CHEMICAL EQUATION: 2 moles H_2 to 1 mole O_2. In this case all reagents are used up.
- One in which we do not have stoichiometric amounts of reagents. This means that the mole ratio of H_2 to O_2 is different than 2:1. In this case, there will be one reagent that will run out first and some of the other reagent will be left over.
- THE REAGENT THAT RUNS OUT FIRST IS KNOWN AS THE LIMITING REAGENT or LIMITING REACTANT.
- In practice, stoichiometric amounts are very hard to be reached because they require perfectly accurate measures.

HOW DO WE DETERMINE THE LR?

There are several ways for determining the limiting reagent! We can, for example, compare the ratio of the number of moles of each reactant that are necessary (as shown by the chemical equation) to the ratio of the number of moles available for the reaction. To do so, we follow the steps below:

- Calculate the number of moles of each reagent.
- Calculate the available mole ratio of the reagent (from the calculated number of moles).
- Determine the required mole ratio (from the chemical equation).
- Compare the available ratio to the required ratio to determine the LR.

Example
How many grams of SO_2 could be formed if 250.0 g FeS_2 react with 200.0 g O_2 according to the equation below?

$$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$$

First, we need to find the L.R:

- Calculate the number of moles of each, FeS_2 and O_2

$$250.0 \text{ g } FeS_2 \times \frac{1 \text{ mole } FeS_2}{119.99 \text{ g } FeS_2} = 2.084 \text{ moles } FeS_2$$

$$200.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 6.250 \text{ moles } O_2$$

- Determine the available mole ratio:

$$\frac{2.084 \text{ mols } FeS_2}{6.250 \text{ mols } O_2} = 0.3334 \text{ mols } FeS_2/\text{mol } O_2$$

- Determine the required ratio:

$$\frac{4 \text{ mols } FeS_2}{11 \text{ moles } O_2} = 0.3636 \text{ mols } FeS_2/\text{mol } O_2$$

- Compare the available and the required ratios:

We have available 0.3334 moles of FeS_2 per mole of O_2. However, the reaction requires 0.3636 moles FeS_2 per mole of O_2. So, we have less moles of FeS_2 per mole of O_2 than what is required and therefore, FeS_2 is the LR.

- Now, we can calculate the maximum number of grams of SO_2 that can be obtained assuming all the FeS_2 reacts.

$$2.084 \text{ mols } FeS_2 \times \frac{8 \text{ mols } SO_2}{4 \text{ mols } FeS_2} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 267.0 \text{ g } SO_2$$

PERCENT YIELD FROM CHEMICAL REACTIONS

In practice, not all chemical reactions go to completion that is the reaction does not actually run until one of the reactant (the limiting reactant) is used up. We need to distinguish in between two yields:

Theoretical Yield: It is the amount of a product that we can potentially obtain if 100% of one reactant was converted to products in a chemical reaction.

Actual Yiel d(also called experimental yied): The amount of products that we can actually obtain after performing the chemical reaction.

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

Example
Suppose we run the reaction as described in the previous example and obtain 200.0 g of SO_2, what would the percent yield be?

The theoretical yield was the one calculated in the previous example: 267.0 g of SO_2.

The actual yield is the one obtained when one actually runs the reaction. It is given to us as 200.0 g of SO_2.

$$\text{The percent yield} = \frac{200.0 \text{ g of } SO_2}{267.0 \text{ g of } SO_2} \times 100 = 74.91\%$$

Example ***(answers will be provided by TA)***
When we react 40.00 g of diboron trioxide, B_2O_3, with 52.50 g of hydrogen fluoride, HF, we isolate 41.65 g of boron trifluoride, BF_3. What is the percent yield of BF_3? The balanced chemical equation is:

$$B_2O_3 + 6HF \longrightarrow 2BF_3 + 3H_2O$$

We first need to calculate the theoretical yield. To do so, we need to determine the LR.

- Calculate number of moles of each reagent:

- Calculate the available ratio:

- Determine the required ratio:

- Compare the available ratio to the required ratio to determine the LR:

- Use the LR to calculate the theoretical yield of BF_3:

- Calculate the percent yield