

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:

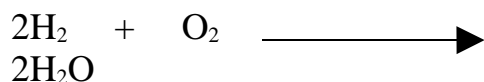
p F90 M.2A, M.2B

M.4, M.6, M.11

p F93

THE LIMITING REAGENT CONCEPT

Consider the reaction illustrated by the following chemical equation:



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.

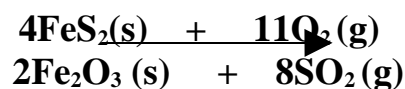
HOW DO WE DETERMINE THE LR?

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 SO₂ could be formed if 250.0 g FeS₂ react with 200.0 g O₂ according to the equation below?



First, we need to find the L.R:

- Calculate the number of moles of each, FeS₂ and O₂

$$\frac{250.0 \text{ g FeS}_2}{119.99 \text{ g FeS}_2} \left| \frac{1 \text{ mole FeS}_2}{119.99 \text{ g FeS}_2} \right. = 2.084 \text{ moles FeS}_2$$

$$\frac{200.0 \text{ g O}_2}{32.00 \text{ g O}_2} \left| \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right. = 6.250 \text{ moles O}_2$$

- Determine the available mole ratio:

$$\frac{2.084 \text{ moles FeS}_2}{6.250 \text{ moles O}_2} = 0.3334 \text{ moles FeS}_2/\text{mol O}_2$$

- Determine the required ratio:

$$\frac{4 \text{ moles FeS}_2}{11 \text{ moles O}_2} = 0.3636 \text{ moles FeS}_2/\text{mol O}_2$$

- Compare the available and the required ratios:

We have available 0.3334 moles of FeS₂ per mole of O₂. However, the reaction requires 0.3636 moles FeS₂ per mole of O₂. So, we have less moles of FeS₂ per mole of O₂ than what is required and therefore, FeS₂ is the LR.

- Now, we can calculate the maximum number of grams of SO₂ that can be obtained assuming all the FeS₂ reacts.

$$\frac{2.084 \text{ moles FeS}_2}{4 \text{ moles FeS}_2} \left| \frac{8 \text{ moles SO}_2}{4 \text{ moles FeS}_2} \right. \left| \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \right. = 267.0 \text{ g SO}_2$$

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 Yield (also called experimental yield): 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 SO₂, what would the percent yield be?

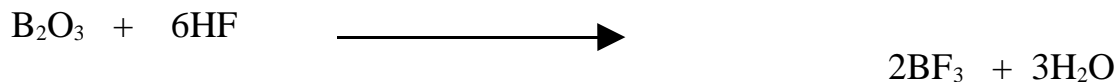
The theoretical yield was the one calculated in the previous example: 267.0 g of SO₂.

The actual yield is the one obtained when one actually runs the reaction. It is given to us as 200.0 g of SO₂.

$$\text{The percent yield} = \frac{200.0 \text{ g of SO}_2}{267.0 \text{ g of SO}_2} \times 100 = 74.91\%$$

Example (*answers will be provided by TA*)

When we react 40.00 g of diboron trioxide, B₂O₃, with 52.50 g of hydrogen fluoride, HF, we isolate 41.65 g of boron trifluoride, BF₃. What is the percent yield of BF₃? The balanced chemical equation is:



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 BF₃:

- Calculate the percent yield