Limiting Reactant Concept

When a chemical reaction is carried out in a laboratory or industrial setting, the reactants are not usually present in the exact molar ratios specified in the balanced chemical equation for the reaction. Most often, on purpose, excess quantities of one or more of the reactants are present.

Numerous reasons exist for having some reactants present in excess. Sometimes such a procedure will cause a reaction to occur more rapidly. For example, large amounts of oxygen make combustible materials burn faster. Sometimes an excess of one reactant will insure that another reactant, perhaps a very expensive one, is completely consumed.

When one or more reactants are present in excess, the excess will not react because there is not enough of the other reactant to react with it. The reactant not in excess, thus, limits the amount of product(s) formed and is called the limiting reactant. The limiting reactant is the reactant in a chemical reaction that determines how much product(s) can be formed.

Let us consider some simple but analogous nonchemical examples of a “limiting reactant” before going to sample limiting reactant calculations. Suppose we have a vending machine that contains forty 25¢ candy bars and that we have 27 quarters. In this case only 27 candy bars may be purchased. The quarters are the limiting reactant. The candy bars are present in excess. Suppose we have ten slices of cheese and eighteen slices of bread and we want to make as many cheese sandwiches as possible using one slice of cheese and two slices of bread per sandwich. The eighteen slices of bread limit us to nine sandwiches; one slice of cheese is left over. The bread is the limiting reactant in this case even though initially more bread (18 slices) than cheese (10 slices) was present. The bread is still limiting because it is used up first.

Now let us proceed to chemical calculations that involve a limiting reactant. Whenever you are given the quantities of two or more reactant in a chemical reaction, it is necessary for you to determine which of the given quantities is the limiting reactant.

Determining the limiting reactant can be accomplished by the following procedure.

1. Determine the number of moles of each of the reactants present.
2. Calculate the number of moles of product then grams of product each of the molar amounts of reactant would produce if it were the only reactant amount given. If more than one product is formed in the reaction, you need to do this mole calculation for only one of the products.
3. The reactant that produces the least number of grams of product is the limiting reactant.

Example:

Silver tarnishes in the presence of hydrogen sulfide (H₂S), a gas that originates from the decay of food, because of the reaction

\[ 4 \text{Ag} + 2 \text{H}_2\text{S} + \text{O}_2 \rightarrow 2 \text{Ag}_2\text{S} + 2 \text{H}_2\text{O} \]
The black product Ag₂S is the “tarnish.” If 25.00 g of Ag, 5.00 g of H₂S, and 4.00 g O₂ are present in a reaction mixture, which is the limiting reactant?

Solution:
To determine the limiting reactant, we determine how many grams of product we can get from each of the reactants. In this particular problem we have two products: Ag₂S and H₂O. It is sufficient to calculate the number of grams of either Ag₂S or H₂O. The decision as to which one is arbitrary; we will choose H₂O.

The calculations will be “grams of reactant to grams of product” calculations. Such calculations will involve the use of the coefficients in the balanced equation. The pathway is

grams reactant -------> moles reactant -------> moles product -------> grams product

Note that we will have to go through a similar calculation three times since we have three reactants: once for Ag, once for H₂S and once for O₂.

For Ag
25.00 g Ag = 2.09 g H₂O

For H₂S
5.00 g H₂S = 2.64 g H₂O

For O₂
4.00 g O₂ = 4.51 g H₂O

The limiting reactant is the reactant that will produce the least number of grams of H₂O. Looking at the numbers just calculated we see that Ag will be the limiting reactant.

Yields: Theoretical, Actual, and Percent

The amount of product isolated in a pure form from a chemical reaction is almost always less than the amount predicted using a limiting reactant calculation. A number of factors contribute to this situation. Some product is almost always lost in the process of its isolation and purification and in such mechanical operations as transferring materials from one container to another. Also, in many reactions side reactions occur that lead to the formation of small amounts of extraneous products. Reactants consumed in these side reactions obviously will not end up in the form of the desired product. The net effect of all of this is that the actual quantities of product isolated are less, sometimes far less, than the theoretically possible amount.

Product loss is specified in terms of percent yield, a measure of the ratio between actual yield and theoretical yield.
Actual yield
Percent Yield = ----------------------- x 100
Theoretical yield

The **actual yield** is the amount of product actually obtained at the end of the experiment. It is always an experimentally determined number; it cannot be calculated. The **theoretical yield** is the maximum amount of product that can be produced from the starting amount of reactants if no losses of any kind occurred. It is always a calculated number obtained from a limiting reactant calculation. It is the smallest amount of product the limiting reactant can produce.

**Example:**

A mixture of 80.0 g of Cr₂O₃ and 8.00 g of C is used to produce elemental Cr by the reaction

\[ \text{Cr}_2\text{O}_3 + 3 \text{C} \rightarrow 2 \text{Cr} + 3 \text{CO} \]

- a) What is the theoretical yield of Cr that can be obtained from the reaction mixture?
- b) The actual yield is 21.7 g Cr. What is the percent yield for the reaction?

**Solution:**

- a) the limiting reactant must be determined before the theoretical yield can be determined. Recalling the procedure for determining the limiting reactant, we calculate the number of grams of Cr that can be produced from each individual reactant using gram-to-gram type calculations. The reactant that produces the smallest grams of product is the limiting reactant and that smallest grams of product is the theoretical yield.

**Grams reactant ----> moles reactant ----> moles product ----> grams product**

For Cr₂O₃

\[ 80.0 \text{ g Cr}_2\text{O}_3 = 54.7 \text{ g Cr} \]

For C

\[ 8.00 \text{ g C} = 23.1 \text{ g Cr} \]

Our calculation show that C is the limiting reagent and the theoretical yield of Cr is 23.1 g.

- b) The percent yield is obtained by dividing the actual yield by the theoretical yield and multiplying by 100.

\[
\% \text{ yield} = \frac{x}{100} = \frac{21.7 \text{ g Cr}}{23.1 \text{ g Cr}} = 93.9 \% \text{ yield}
\]