

# General Chemistry 1

## Quarter 1, Lesson 14

### LIMITING AND EXCESS REAGENTS

#### Objectives:

- Define the concept of limiting and excess reactants.
- Determine the limiting and excess reactants using mole ratios from balanced chemical equations.

When a chemical reaction is carried out in a flask, the amount of reactants are not always present in the exact proportion stoichiometrically determined from the balanced equation. To make sure that the more expensive reagent is completely used up and converted to the desired product, chemists usually add the cheaper reagent in excess quantity. The reagent that has completely reacted and used up in a reaction is called the **limiting reagent**. The **excess reagent** is the reactant that is present in quantity higher than what is required to react with the limiting reagent.

**Sample Problem 1.** Let's use the chemical equation below:



Ammonia,  $NH_3$ , is synthesized from the reaction of  $H_2$  and  $N_2$  gases. Suppose 6 moles of  $H_2$  was initially mixed with 4 moles of  $N_2$  gas at high pressure. To determine which of the 2 reactants is the limiting reagent, the amount of  $NH_3$  produced must be computed given the number of moles of  $H_2$  and  $N_2$  and the mole-mole factor from the balanced equation.

$$\text{moles } NH_3 = \cancel{\# \text{ moles of } H_2} \times \frac{2 \text{ moles } NH_3}{\cancel{3 \text{ moles } H_2}}$$

$$\begin{aligned} \text{moles } NH_3 &= 6 \cancel{\text{ moles of } H_2} \times \frac{2 \text{ moles } NH_3}{\cancel{3 \text{ moles } H_2}} \\ \text{moles } NH_3 &= 4 \text{ moles} \end{aligned}$$

$$\text{moles } NH_3 = \cancel{\text{moles of } N_2} \times \frac{2 \text{ moles } NH_3}{\cancel{1 \text{ mole } N_2}}$$

$$\begin{aligned} \text{moles } NH_3 &= 4 \cancel{\text{ moles of } N_2} \times \frac{2 \text{ moles } NH_3}{\cancel{1 \text{ mole } N_2}} \\ \text{moles } NH_3 &= 8 \text{ moles} \end{aligned}$$

From the computed values we determined that if 6 moles of  $H_2$  completely reacts with  $N_2$ , it can produce 4 moles of  $NH_3$  while 4 moles of  $N_2$  can produce 8 moles of  $NH_3$  when fully used up. Since there is less amount of  $NH_3$  produced with 6 moles of  $H_2$  than 4 moles of  $N_2$ ,  $H_2$  gas is the limiting reagent while the  $N_2$  gas is the reagent in excess. To determine how much of the 4 moles of  $N_2$  is in excess, we will use the mole-mole factor of  $N_2$  and  $H_2$ .

$$\begin{aligned} \text{moles } N_2 &= 6 \cancel{\text{ moles of } H_2} \times \frac{1 \text{ mole } N_2}{\cancel{3 \text{ moles } H_2}} \\ \text{moles } N_2 &= 2 \text{ moles} \end{aligned}$$

The number of moles of  $N_2$  required to fully react to 6 moles of  $H_2$  is only 2 moles. Thus, the initial 4 moles of  $N_2$  has an excess of 2 moles.

$$\text{excess } N_2 = 2 \text{ moles} - 2 \text{ moles} = 2 \text{ moles}$$

**Sample Problem 2.** The combustion of ethane produces carbon dioxide and water shown in the reaction below:



1. How many moles of  $CO_2$  are produced with 56.2 moles  $C_2H_6$  and 73.4 moles of  $O_2$ ?  
Solution:

$$\begin{aligned} \text{moles } CO_2 &= 56.2 \cancel{\text{ moles of } C_2H_6} \times \frac{4 \text{ moles } CO_2}{\cancel{2 \text{ moles } C_2H_6}} \\ \text{moles } CO_2 &= 112.4 \text{ moles} \end{aligned}$$

$$\begin{aligned} \text{moles } CO_2 &= 73.4 \cancel{\text{ moles of } O_2} \times \frac{4 \text{ moles } CO_2}{\cancel{7 \text{ moles } O_2}} \\ \text{moles } CO_2 &= 41.9 \text{ moles} \end{aligned}$$

2. Which reagent is the limiting reagent? Which is the excess reagent?  
*The Limiting reagent is  $O_2$  while the excess reagent is  $C_2H_6$ .*
3. How many moles of  $C_2H_6$  are in excess?

# Sample Problem Of Limiting Reactants Answer Key

**RD Boyd**



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