Ch. 9.4 Lecture Guide

H. Chemistry

9.4

- **Review:**
  - Define stoichiometry.

The process of using a chemical equation to calculate the relative masses of reactants and products in a reaction.

- **Stoichiometric quantities:**
  - A stoichiometric quantity uses exactly the correct amounts of each reactant so that both "run out" or are used up in a chemical reaction.

  - **Question: Why would chemists prefer to use stoichiometric quantities?**
    - If reactants were not mixed in stoichiometric quantities, some reactant will not be completely used up and this is very wasteful.

  - **Consider the following chemical reaction:** $CH_4\ (g) + H_2O\ (g) \rightarrow 3H_2\ (g) + CO_2\ (g)$
    - What mass of water is required to react exactly with 249 g of methane?

$$249\,g\ \cancel{CH_4} \times \frac{1\ \cancel{mol\ CH_4}}{16.04\,\cancel{g\ CH_4}} \times \frac{1\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ CH_4}} \times \frac{18.02\,g\ H_2O}{1\ \cancel{mol\ H_2O}} = \boxed{280.\,g\ H_2O}$$

Start with the starting unit

So, 280. g of H_2O will react exactly with 249 g CH_4.

- What would happen if we reacted 300 g of water with 249 g of CH_4?

There would be left-over water. Chemist would say that water would be in excess.

- In this example, methane is the **limiting reactant**, or **limiting reagent**, because it limits the amount of product that will be created.

- In chemistry, it is important to know which reactant is the limiting reactant in order to predict that amount of product that will form.

- **Identifying the Limiting Reagent**
  - **Consider the following problem:** Suppose 2.50×10^4 g of nitrogen gas (N_2) reacts with 5.00×10^3 g of hydrogen gas (H_2) to produce ammonia (NH_3). Calculate the mass of ammonia produced when this reaction is run to completion. **Note: This problem is different from others that we have done because now we are mixing two specified amounts of reactants together.**

Review

- **Step 1:** Write the balanced equation.

$$N_2(g) + \underline{3}H_2(g) \longrightarrow \underline{2}NH_3(g)$$

Review

- **Step 2:** Convert the known masses of the reactants from grams to moles.

$$2.50 \times 10^4\ \cancel{g\ N_2} \times \frac{1\ mol\ N_2}{28.02\ \cancel{g\ N_2}} = 8.92 \times 10^2\ mol\ N_2$$

$$5.00 \times 10^3\ \cancel{g\ H_2} \times \frac{1\ mol\ H_2}{2.016\ \cancel{g\ H_2}} = 2.48 \times 10^3\ mol\ H_2$$

New

- **Step 3:** Use mole ratios to determine which factor is limiting.

Calculate the amount of moles of H_2 required to react exactly with the available amount of N_2.

$$\underbrace{8.92 \times 10^2\ mol\ \cancel{N_2}}_{available} \times \frac{3\ mol\ H_2}{1\ mol\ \cancel{N_2}} = \underbrace{2.68 \times 10^3\ mol\ H_2}_{required}$$

Calculate the amount of moles of N_2 required to react exactly with the available amount of H_2.

$$\underbrace{2.48 \times 10^3\ mol\ \cancel{H_2}}_{available} \times \frac{1\ mol\ N_2}{3\ mol\ \cancel{H_2}} = \underbrace{8.27 \times 10^2\ mol\ N_2}_{required}$$

$$\underbrace{8.92 \times 10^2\ mol\ N_2}_{available} > \underbrace{8.27 \times 10^2\ mol\ N_2}_{required}$$

$$\underbrace{2.48 \times 10^3\ mol\ H_2}_{available} < \underbrace{2.68 \times 10^3}_{required}$$

★ Therefore, H_2 is the limiting reactant because there is more required than is available.

Review

- **Step 4:** Use moles of limiting reactant available and mole ratio to calculate the number of moles of product formed.

$$2.48 \times 10^3 \text{ mol } \cancel{H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } \cancel{H_2}} = 1.65 \times 10^3 \text{ mol } NH_3$$

Review

- **Step 5:** Convert moles of product to grams of product.

$$1.65 \times 10^3 \text{ mol } \cancel{NH_3} \times \frac{17.03 g \; NH_3}{1 \text{ mol } \cancel{NH_3}} = \boxed{2.81 \times 10^4 g \; NH_3}$$