

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: $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow 3\text{H}_2(\text{g}) + \text{CO}_2(\text{g})$

- What mass of water is required to react exactly with 249 g of methane?

Start with the starting unit →

$$249 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 280. \text{ g H}_2\text{O}$$

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.

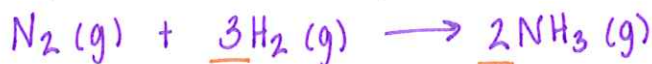
- In chemistry, it is important to know which reactant is the limiting reactant in order to predict that amount of product that will form.

➤ Identifying the Limiting Reagent

- Consider the following problem: Suppose 2.50×10^4 g of N_2 (nitrogen) gas reacts with 5.00×10^3 g of H_2 (hydrogen) gas 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.



Review

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

$$2.50 \times 10^4 \text{ g } \text{N}_2 \times \frac{1 \text{ mol } \text{N}_2}{28.02 \text{ g } \text{N}_2} = 8.92 \times 10^2 \text{ mol } \text{N}_2$$

$$5.00 \times 10^3 \text{ g } \text{H}_2 \times \frac{1 \text{ mol } \text{H}_2}{2.016 \text{ g } \text{H}_2} = 2.48 \times 10^3 \text{ mol } \text{H}_2$$

New

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

$$8.92 \times 10^2 \text{ mol } \text{N}_2 \text{ available} \times \frac{3 \text{ mol } \text{H}_2}{1 \text{ mol } \text{N}_2} = 2.68 \times 10^3 \text{ mol } \text{H}_2 \text{ required}$$

$$2.48 \times 10^3 \text{ mol } \text{H}_2 \text{ available} \times \frac{1 \text{ mol } \text{N}_2}{3 \text{ mol } \text{H}_2} = 8.27 \times 10^2 \text{ mol } \text{N}_2 \text{ required}$$

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

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

$$8.92 \times 10^2 \text{ mol } \text{N}_2 \text{ available} > 8.27 \times 10^2 \text{ mol } \text{N}_2 \text{ required}$$

$$2.48 \times 10^3 \text{ mol } \text{H}_2 \text{ available} < 2.68 \times 10^3 \text{ mol } \text{H}_2 \text{ required}$$

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.

$$2.48 \times 10^3 \text{ mol } \cancel{\text{H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol } \cancel{\text{H}_2}} = 1.65 \times 10^3 \text{ mol NH}_3$$

Review

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

$$1.65 \times 10^3 \text{ mol } \cancel{\text{NH}_3} \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol } \cancel{\text{NH}_3}} = \boxed{2.81 \times 10^4 \text{ g NH}_3}$$