ChemTalk

Unit 3: Stoichiometry

Stoichiometry: Limiting Reagents and Percent Yield

Stoichiometry is a pivotal concept that forms the foundation for the understanding of chemical reactions. Identifying **limiting & excess reactants** is crucial in determining the maximum product formed in a chemical reaction, highlighting the significance of balanced equations. Reactants that are left over after the reaction has completed are excess reactants. **Theoretical yield**, a key quantitative factor, represents the maximum product achievable under ideal conditions, often calculated using stoichiometric coefficients. **Percent yield**, expressed as a percentage, gauges reaction efficiency by comparing actual yield to theoretical yield. These concepts are vital in chemical calculations, providing valuable insights into process efficiency and outcomes.

1. A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen by mass. What is the empirical formula of this compound?

2. Consider the reaction: $2H_2 + O_2 \rightarrow 2H_2O$. If you have 6 moles of H_2 and 2 moles of O_2, which is the limiting reactant, and what is the maximum moles of H_2O that can be produced?

3. In a chemical reaction, 4 moles of reactant A react with 6 moles of reactant B to produce 12 moles of product C, according to the balanced equation: $4A + 6B \rightarrow 12C$. What is the theoretical yield of C when you have 10 moles of A and 6 moles of B available for the reaction?

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

4. You observe the above combustion reaction. In this experiment, you had 57.0 grams of CH_4 and 102.0 grams of O_2. However, because of an accident caused in the lab, you are now expecting to produce 50 grams of H_2O. However, you only obtained 40 grams of H_2O. Calculate the percent yield for this reaction.

5. You are conducting a chemical reaction between hydrogen gas (H_2) and nitrogen gas (N_2) according to the balanced equation:

$$3\ H_2 + N_2 \rightarrow 2\ NH_3$$

You have 8 moles of H_2 and 5 moles of N_2 available for the reaction.

a. Which reactant is the limiting reactant?

b. If the actual yield of ammonia (NH_3) obtained from this reaction is 10 grams, calculate the percent yield.

6. Consider the combustion reaction of ethane (C_2H_6):

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

a. If you begin with 10.0 grams of ethane and 30.0 grams of oxygen, determine the limiting reagent and calculate the mass of the excess reagent remaining after the reaction is complete.

b. In a different experiment, starting with 15.0 grams of ethane and 20.0 grams of oxygen, find the limiting reagent and calculate the mass of each product formed.

i. If you are expecting to form 50.0 grams of H_2O, what is the percent yield?

7. Consider the reaction between iron(III) oxide (Fe_2O_3) and carbon monoxide (CO) to produce iron (Fe) and carbon dioxide (CO_2):

$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

a. If you start with 15.0 grams of Fe_2O_3 and 12.0 grams of CO, determine the limiting reagent and calculate the mass of each product formed.

i. If you form 6.39 g of Fe, what is the percent yield?

b. In a separate experiment, starting with 25.0 grams of Fe_2O_3 and 18.0 grams of CO, find the limiting reagent and calculate the mass of the excess reagent remaining after the reaction is complete.

ANSWER KEY

1. A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen by mass. What is the empirical formula of this compound?

   Step 1: Assume a 100g sample.

   Step 2: Convert percentages to grams.
   Carbon: 40 % → 40 g
   Hydrogen: 6.7 % → 6.7 g
   Oxygen: 53.3 % → 53.3 g

   Step 3: Find moles of each element. These can be found on the periodic table. Use the molar mass of each element to find the moles:
   Moles of Carbon (C): 40g÷12.01g/mol ≈ **3.33 mol C**
   Moles of Hydrogen (H): 6.7g÷1.01g/mol ≈ **6.64 mol H**
   Moles of Oxygen (O): 53.3 g÷16.00 g/mol ≈ **3.33 mol O**

   Step 4: Find the simplest mole ratio. Divide each mole value by the smallest number of moles among the elements (which is 3.33):
   C : 3.33 mol ÷ 3.33 mol ≈ **1**
   H : 6.64 mol ÷ 3.33 mol ≈ **2**
   O : 3.33 mol ÷ 3.33 mol ≈ **1**

   Step 5: Write the empirical formula. The empirical formula is **CH_2O**.

2. Consider the reaction: $2H_2 + O_2 \rightarrow 2H_2O$. If you have 6 moles of H_2 and 2 moles of O_2, which is the limiting reactant, and what is the maximum moles of H_2O that can be produced?

   To find the limiting reactant, we need to calculate the moles of H_2O that can be produced from both reactants. From 6 moles of H_2, you can produce 6 moles of H_2O (2:2 mole ratio). From 2 moles of O_2, you can produce 4 moles of H_2O (1:2 mole ratio). Since O_2 limits the formation of H_2O to 4 moles, it is the limiting reactant. The maximum moles of H_2O that can be produced is 4 moles.

3. In a chemical reaction, 4 moles of reactant A react with 6 moles of reactant B to produce 12 moles of product C, according to the balanced equation: 4A + 6B → 12C. What is the theoretical yield of C when you have 10 moles of A and 6 moles of B available for the reaction?

Given the balanced equation: 4A + 6B → 12C

Available moles:
Moles of A= 10
Moles of B= 6

Using the mole ratio from the equation, we can determine how many moles of C are produced from each amount of A and B.

$$10\ moles\ A \times \frac{12\ moles\ C}{4\ moles\ A} = 30\ moles\ of\ C$$

$$6\ moles\ B \times \frac{12\ moles\ C}{6\ moles\ B} = 12\ moles\ of\ C$$

As we can see from the above equations, B is the limiting reagent as 6 moles of B can only produce 12 moles of C. Thus, the theoretical yield of C is 12 moles and reactant A would be in excess.

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

4. You observe the above combustion reaction. In this experiment, you had 57.0 grams of CH_4 and 102.0 grams of O_2. However, because of an accident caused in the lab, you are now expecting to produce 50 grams of H_2O. However, you only obtained 40 grams of H_2O. Calculate the percent yield for this reaction.

   To calculate percent yield, use the formula:

   $$\frac{Actual\ Yield}{Theoretical\ Yield} \times 100\% = Percent\ Yield$$

   In this case, the actual yield is 40 grams, and the theoretical yield was expected to be 50 grams. So, the percent yield is:

   $$\frac{40\ g}{50\ g} \times 100\% = 80\%$$

5. You are conducting a chemical reaction between hydrogen gas (H_2) and nitrogen gas (N_2) according to the balanced equation:

   $$3H_2 + N_2 \rightarrow 2NH_3$$

   You have 8 moles of H_2 and 5 moles of N_2 available for the reaction.

   a. Which reactant is the limiting reactant?

b. If the actual yield of ammonia (NH_3) obtained from this reaction is 82 grams, calculate the percent yield.

a) To find the limiting reactant, we need to compare calculate how much product each reactant alone can produce.

For H_2:

$$8\ moles\ H_2 \times \frac{2\ mol\ NH_3}{3\ mol\ H_2} = 5.33\ mol\ NH_3$$

For N_2:

$$5\ moles\ N_2 \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} = 10\ mol\ NH_3$$

Because H_2 produces fewer moles of NH_3, it is the limiting reactant.

b) Now, let's calculate the theoretical yield of NH_3 based on the limiting reactant, which is H_2. We can use the number of moles from part a to determine the theoretical yield of NH_3. The molar mass of NH_3 is approximately 17.03 g/mol.

$$Theoretical\ Yield\ of\ NH_3 = 5.33\ mol\ NH_3 \times \frac{17.03\ g\ NH_3}{1\ mol\ NH_3} = 90.83\ g\ NH_3$$

If we only yielded 82.0 grams of NH_3, then we can use the percent yield formula.

$$Percent\ Yield = \frac{(82.0\ g\ of\ NH_3)}{(90.83\ g\ of\ NH_3)} \times 100\% = 90.3\%\ Yield$$

6. Consider the combustion reaction of ethane (C_2H_6):

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

   a. If you begin with 10.0 grams of ethane and 30.0 grams of oxygen, determine the limiting reagent.
   b. In a different experiment, starting with 15.0 grams of ethane and 20.0 grams of oxygen, find the limiting reagent and calculate the mass of each product formed.
      i. If you are expecting to form 17.0 grams of H_2O, what is the percent yield?

a) **Given:**
Initial mass of C_2H_6 = 10.0 grams

Initial moles of O_2 = 30.0 grams

First we need to find the number of moles for each substance:

$$Moles\ of\ C_2H_6 = 10\ g\ C_2H_6 \times \frac{1\ mol\ C_2H_6}{30.07\ g\ C_2H_6} = 0.333\ mol\ C_2H_6$$

$$Moles\ of\ O_2 = 30\ g\ O_2 \times \frac{1\ mol\ O_2}{32.00\ g} = 0.938\ mol\ O_2$$

Next, we can convert each of these values to moles of H_2O to determine the limiting reagents.

C_2H_6:

$$0.333\ mol\ C_2H_6 \times \frac{6\ mol\ H_2O}{2\ mol\ C_2H_6} = 0.998\ mol\ H_2O$$

O_2:

$$0.938\ mol\ O_2 \times \frac{6\ mol\ H_2O}{7\ mol\ O_2} = 0.804\ mol\ H_2O$$

Given that oxygen produces less moles of H_2O, it is the limiting reagent.

b) **Calculate moles of C_2H_6 and O_2 using their molar masses:**

$$Moles\ of\ C_2H_6 = 15\ g\ C_2H_6 \times \frac{1\ mol\ C_2H_6}{30.07\ g\ C_2H_6} = 0.499\ mol\ C_2H_6$$

$$Moles\ of\ O_2 = 20\ g\ O_2 \times \frac{1\ mol\ O_2}{32.00\ g} = 0.625 mol\ O_2$$

Determine the limiting reagent:

C_2H_6:

$$0.499\ mol\ C_2H_6 \times \frac{6\ mol\ H_2O}{2\ mol\ C_2H_6} = 1.497\ mol\ H_2O$$

O_2:

$$0.625\ mol\ O_2 \times \frac{6\ mol\ H_2O}{7\ mol\ O_2} = 0.536\ mol\ H_2O$$

O_2 is the limiting reactant because it produces less moles of H_2O.

Calculate the mass of the products using O_2 as the limiting reactant:

Using the moles of H_2O from part a:

$$Grams\ of\ H_2O\ formed = 0.536\ mol\ H_2O \times \frac{18.02\ g\ H_2O}{1\ mol\ H_2O} = 9.66\ g\ H_2O$$

$$Grams\ of\ CO_2\ formed:\ 0.625\ mol\ O_2 \times \frac{4\ mol\ CO_2}{7\ mol\ O_2} \times \frac{44.01\ g\ CO_2}{1\ mol\ CO_2} = 15.7\ g\ CO_2$$

i) If you expected to form 17.0 grams of H_2O, use the equation:

$$Percent\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100\% = \frac{9.66\ g\ of\ H_2O}{17.0\ g\ of\ H_2O} \times 100\% = 56.8\%\ Yield$$

7. Consider the reaction between iron(III) oxide (Fe_2O_3) and carbon monoxide (CO) to produce iron (Fe) and carbon dioxide (CO_2):

$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

   a. If you start with 15.0 grams of Fe_2O_3 and 12.0 grams of CO, determine the limiting reagent and calculate the mass of each product formed.
      i. If you form 6.39 grams of Fe, what is the percent yield?
   b. In a separate experiment, starting with 25.0 grams of Fe_2O_3 and 18.0 grams of CO, find the limiting reagent and calculate the mass of the excess reagent remaining after the reaction is complete.

a. **Calculate moles of one of the products from both reactants:**

$$Moles\ of\ Fe_2O_3\ = 15\ g\ Fe_2O_3 \times \frac{1\ mol\ Fe_2O_3}{159.69\ g\ Fe_2O_3} = 0.094\ mol\ Fe_2O_3 \times \frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3} = 0.187\ mol\ Fe$$

$$Moles\ of\ CO\ = 12\ g\ CO \times \frac{1\ mol\ CO}{28.01\ g\ CO} = 0.428\ mol\ CO \times \frac{2\ mol\ Fe}{3\ mol\ CO} = 0.286\ mol\ Fe$$

As Fe_2O_3 produces less moles of Fe, it is the limiting reagent. Given this, we can calculate the mass of Fe and CO_2 produced from moles of Fe_2O_3.

We have moles of Fe from the calculation above:

$$Grams\ of\ Fe\ formed\ = 0.187\ mol\ Fe \times \frac{55.85\ g\ Fe}{1\ mol\ Fe} = 10.4\ g\ Fe$$

For CO_2, we can start from moles of Fe_2O_3:

$$Grams\ of\ CO_2\ formed\ = 0.094\ mol\ Fe_2O_3 \times \frac{3\ mol\ CO_2}{1\ mol\ Fe_2O_3} \times \frac{44.01 g\ CO_2}{1\ mol\ CO_2} = 12.4 g\ CO_2$$

i. If we were expecting to produce 10.4 g of Fe but only produced 6.39 g, the percent yield would be:

$$Percent\ Yield\ = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100\% = \frac{6.39\ g\ of\ Fe}{10.4\ g\ Fe} \times 100\% = 61.4\ \%\ Yield$$

b. **Calculate moles of one of the products from both reactants:**

$$Moles\ of\ Fe_2O_3\ = 25\ g\ Fe_2O_3 \times \frac{1\ mol\ Fe_2O_3}{159.69\ g\ Fe_2O_3} = 0.157\ mol\ Fe_2O_3 \times \frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3} = 0.313\ mol\ Fe$$

$$Moles\ of\ CO\ = 18\ g\ CO \times \frac{1\ mol\ CO}{28.01\ g\ CO} = 0.643\ mol\ CO \times \frac{2\ mol\ Fe}{3\ mol\ CO} = 0.428\ mol\ Fe$$

Fe_2O_3 is the limiting reagent as it produces less product.

Calculate unreacted amounts:
No Fe_2O_3 will remain as it will all react. Thus, we can calculate the remaining mass for CO from the reacted moles of Fe_2O_3.

$$Grams\ of\ CO\ reacted = 0.157\ mol\ Fe_2O_3 \times \frac{3\ mol\ CO}{1\ mol\ Fe_2O_3} \times \frac{28.01\ g\ CO}{1\ mol\ CO} = 13.19\ g\ CO\ reacted$$

$$Grams\ of\ CO\ in\ excess = 18.0\ g\ CO - 13.19\ g\ CO\ reacted = 4.81\ g\ CO\ excess$$