



### Unit 3: Stoichiometry

#### Stoichiometry: Limiting Reagents and Percent Yield

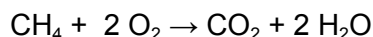
Stoichiometry is a pivotal concept that forms the foundation for the understanding of chemical reactions. Identifying **limiting & excess reactants** is crucial in determining the maximum product formed in a chemical reaction, highlighting the significance of balanced equations. Reactants that are left 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:  $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ . If you have 6 moles of  $\text{H}_2$  and 2 moles of  $\text{O}_2$ , which is the limiting reactant, and what is the maximum moles of  $\text{H}_2\text{O}$  that can be produced?

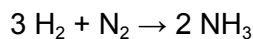
3. In a chemical reaction, 4 moles of reactant A react with 6 moles of reactant B to produce 12 moles of product C, according to the balanced equation:  $4\text{A} + 6\text{B} \rightarrow 12\text{C}$ . What is the theoretical yield of C when you have 10 moles of A and 6 moles of B available for the reaction?

## Unit 3: Stoichiometry



4. You observe the above combustion reaction. In this experiment, you had 57.0 grams of  $\text{CH}_4$  and 102.0 grams of  $\text{O}_2$ . However, because of an accident caused in the lab, you are now expecting to produce 50 grams of  $\text{H}_2\text{O}$ . However, you only obtained 40 grams of  $\text{H}_2\text{O}$ . Calculate the percent yield for this reaction.

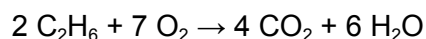
5. You are conducting a chemical reaction between hydrogen gas ( $\text{H}_2$ ) and nitrogen gas ( $\text{N}_2$ ) according to the balanced equation:



You have 8 moles of  $\text{H}_2$  and 5 moles of  $\text{N}_2$  available for the reaction.

- a. Which reactant is the limiting reactant?
- b. If the actual yield of ammonia ( $\text{NH}_3$ ) obtained from this reaction is 10 grams, calculate the percent yield.

6. Consider the combustion reaction of ethane ( $\text{C}_2\text{H}_6$ ):



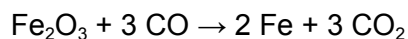
- a. If you begin with 10.0 grams of ethane and 30.0 grams of oxygen, determine the limiting reagent and calculate the mass of the excess reagent remaining after the reaction is complete.



### Unit 3: Stoichiometry

- 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  $\text{H}_2\text{O}$ , what is the percent yield?

7. Consider the reaction between iron(III) oxide ( $\text{Fe}_2\text{O}_3$ ) and carbon monoxide ( $\text{CO}$ ) to produce iron ( $\text{Fe}$ ) and carbon dioxide ( $\text{CO}_2$ ):



- a. If you start with 15.0 grams of  $\text{Fe}_2\text{O}_3$  and 12.0 grams of  $\text{CO}$ , determine the limiting reagent and calculate the mass of each product formed.
- i. If you form 6.39 g of  $\text{Fe}$ , what is the percent yield?
- b. In a separate experiment, starting with 25.0 grams of  $\text{Fe}_2\text{O}_3$  and 18.0 grams of  $\text{CO}$ , 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):  $40\text{g} \div 12.01\text{g/mol} \approx \mathbf{3.33 \text{ mol C}}$

Moles of Hydrogen (H):  $6.7\text{g} \div 1.01\text{g/mol} \approx \mathbf{6.64 \text{ mol H}}$

Moles of Oxygen (O):  $53.3 \text{ g} \div 16.00 \text{ g/mol} \approx \mathbf{3.33 \text{ mol O}}$

**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 :  $3.33 \text{ mol} \div 3.33 \text{ mol} \approx \mathbf{1}$

H :  $6.64 \text{ mol} \div 3.33 \text{ mol} \approx \mathbf{2}$

O :  $3.33 \text{ mol} \div 3.33 \text{ mol} \approx \mathbf{1}$

**Step 5:** Write the empirical formula. The empirical formula is **CH<sub>2</sub>O**.

2. Consider the reaction:  $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ . If you have 6 moles of H<sub>2</sub> and 2 moles of O<sub>2</sub>, which is the limiting reactant, and what is the maximum moles of H<sub>2</sub>O that can be produced?

To find the limiting reactant, we need to calculate the moles of H<sub>2</sub>O that can be produced from both reactants. From 6 moles of H<sub>2</sub>, you can produce 6 moles of H<sub>2</sub>O (2:2 mole ratio). From 2 moles of O<sub>2</sub>, you can produce 4 moles of H<sub>2</sub>O (1:2 mole ratio). Since O<sub>2</sub> limits the formation of H<sub>2</sub>O to 4 moles, it is the limiting reactant. The maximum moles of H<sub>2</sub>O 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:  $4\text{A} + 6\text{B} \rightarrow 12\text{C}$ . 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:  $4A + 6B \rightarrow 12C$

Available moles:

Moles of A = 10

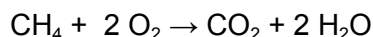
Moles of B = 6

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

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

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

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.



4. You observe the above combustion reaction. In this experiment, you had 57.0 grams of  $\text{CH}_4$  and 102.0 grams of  $\text{O}_2$ . However, because of an accident caused in the lab, you are now expecting to produce 50 grams of  $\text{H}_2\text{O}$ . However, you only obtained 40 grams of  $\text{H}_2\text{O}$ . Calculate the percent yield for this reaction.

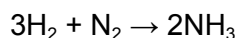
To calculate percent yield, use the formula:

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

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

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

5. You are conducting a chemical reaction between hydrogen gas ( $\text{H}_2$ ) and nitrogen gas ( $\text{N}_2$ ) according to the balanced equation:



You have 8 moles of  $\text{H}_2$  and 5 moles of  $\text{N}_2$  available for the reaction.

- a. Which reactant is the limiting reactant?

## Unit 3: Stoichiometry

b. If the actual yield of ammonia ( $\text{NH}_3$ ) 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  $\text{H}_2$ :**

$$8 \text{ moles } \text{H}_2 \times \frac{2 \text{ mol } \text{NH}_3}{3 \text{ mol } \text{H}_2} = 5.33 \text{ mol } \text{NH}_3$$

**For  $\text{N}_2$ :**

$$5 \text{ moles } \text{N}_2 \times \frac{2 \text{ mol } \text{NH}_3}{1 \text{ mol } \text{N}_2} = 10 \text{ mol } \text{NH}_3$$

Because  $\text{H}_2$  produces fewer moles of  $\text{NH}_3$ , it is the limiting reactant.

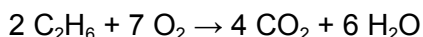
- b) Now, let's calculate the theoretical yield of  $\text{NH}_3$  based on the limiting reactant, which is  $\text{H}_2$ . We can use the number of moles from part a to determine the theoretical yield of  $\text{NH}_3$ . The molar mass of  $\text{NH}_3$  is approximately 17.03 g/mol.

$$\text{Theoretical Yield of } \text{NH}_3 = 5.33 \text{ mol } \text{NH}_3 \times \frac{17.03 \text{ g } \text{NH}_3}{1 \text{ mol } \text{NH}_3} = 90.83 \text{ g } \text{NH}_3$$

If we only yielded 82.0 grams of  $\text{NH}_3$ , then we can use the percent yield formula.

$$\text{Percent Yield} = \frac{(82.0 \text{ g of } \text{NH}_3)}{(90.83 \text{ g of } \text{NH}_3)} \times 100\% = 90.3\% \text{ Yield}$$

6. Consider the combustion reaction of ethane ( $\text{C}_2\text{H}_6$ ):



- If you begin with 10.0 grams of ethane and 30.0 grams of oxygen, determine the limiting reagent.
- 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.
  - If you are expecting to form 17.0 grams of  $\text{H}_2\text{O}$ , what is the percent yield?

- a) **Given:**

Initial mass of  $\text{C}_2\text{H}_6 = 10.0$  grams

### Unit 3: Stoichiometry

Initial moles of  $O_2 = 30.0$  grams

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

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

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

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

$C_2H_6$ :

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

$O_2$ :

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

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:**

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

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

**Determine the limiting reagent:**

$C_2H_6$ :

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

$O_2$ :

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

**$O_2$  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:

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

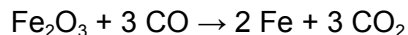
$$\text{Grams of } CO_2 \text{ formed: } 0.625 \text{ mol } O_2 \times \frac{4 \text{ mol } CO_2}{7 \text{ mol } O_2} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 15.7 \text{ g } CO_2$$

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

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

## Unit 3: Stoichiometry

7. Consider the reaction between iron(III) oxide ( $\text{Fe}_2\text{O}_3$ ) and carbon monoxide ( $\text{CO}$ ) to produce iron ( $\text{Fe}$ ) and carbon dioxide ( $\text{CO}_2$ ):



- a. If you start with 15.0 grams of  $\text{Fe}_2\text{O}_3$  and 12.0 grams of  $\text{CO}$ , determine the limiting reagent and calculate the mass of each product formed.
- i. If you form 6.39 grams of  $\text{Fe}$ , what is the percent yield?
- b. In a separate experiment, starting with 25.0 grams of  $\text{Fe}_2\text{O}_3$  and 18.0 grams of  $\text{CO}$ , 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:**

$$\text{Moles of } \text{Fe}_2\text{O}_3 = 15 \text{ g } \text{Fe}_2\text{O}_3 \times \frac{1 \text{ mol } \text{Fe}_2\text{O}_3}{159.69 \text{ g } \text{Fe}_2\text{O}_3} = 0.094 \text{ mol } \text{Fe}_2\text{O}_3 \times \frac{2 \text{ mol } \text{Fe}}{1 \text{ mol } \text{Fe}_2\text{O}_3} = 0.187 \text{ mol } \text{Fe}$$

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

As  $\text{Fe}_2\text{O}_3$  produces less moles of  $\text{Fe}$ , it is the limiting reagent. Given this, we can calculate the mass of  $\text{Fe}$  and  $\text{CO}_2$  produced from moles of  $\text{Fe}_2\text{O}_3$ .

We have moles of  $\text{Fe}$  from the calculation above:

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

For  $\text{CO}_2$ , we can start from moles of  $\text{Fe}_2\text{O}_3$ :

$$\text{Grams of } \text{CO}_2 \text{ formed} = 0.094 \text{ mol } \text{Fe}_2\text{O}_3 \times \frac{3 \text{ mol } \text{CO}_2}{1 \text{ mol } \text{Fe}_2\text{O}_3} \times \frac{44.01 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} = 12.4 \text{ g } \text{CO}_2$$

- i. If we were expecting to produce 10.4 g of  $\text{Fe}$  but only produced 6.39 g, the percent yield would be:

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

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

$$\text{Moles of } \text{Fe}_2\text{O}_3 = 25 \text{ g } \text{Fe}_2\text{O}_3 \times \frac{1 \text{ mol } \text{Fe}_2\text{O}_3}{159.69 \text{ g } \text{Fe}_2\text{O}_3} = 0.157 \text{ mol } \text{Fe}_2\text{O}_3 \times \frac{2 \text{ mol } \text{Fe}}{1 \text{ mol } \text{Fe}_2\text{O}_3} = 0.313 \text{ mol } \text{Fe}$$

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

**$\text{Fe}_2\text{O}_3$  is the limiting reagent as it produces less product.**

**Calculate unreacted amounts:**

No  $\text{Fe}_2\text{O}_3$  will remain as it will all react. Thus, we can calculate the remaining mass for  $\text{CO}$  from the reacted moles of  $\text{Fe}_2\text{O}_3$ .





### Unit 3: Stoichiometry

$$\text{Grams of CO reacted} = 0.157 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{28.01 \text{ g CO}}{1 \text{ mol CO}} = 13.19 \text{ g CO reacted}$$

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