Mass Balance

Week 1
Stoichiometry and Chemical Reactions

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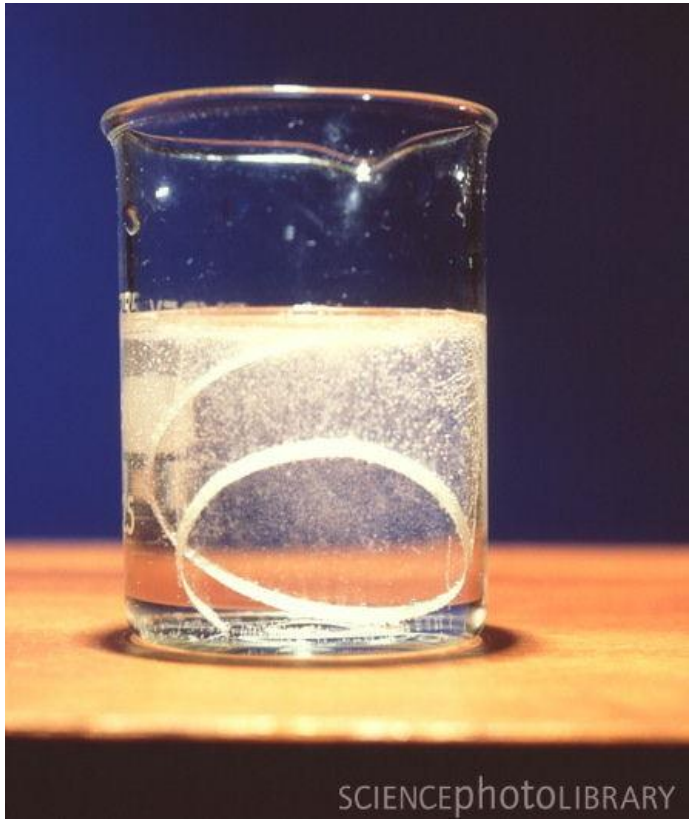
Writing Chemical Equations

- Represents a chemical reaction where reactants transform into products.
- General format: Reactants ⟶ Products.
- Symbols for states: (s) for solids, (l) for liquids, (g) for gases, (aq) for aqueous solutions.

Objectives:

- Recognize signs of chemical reactions.
- Recognize the reactants and products in a reaction.
- Identify different types of chemical reactions.

Signs of Chemical Reactions



Evolution of a gas

Signs of Chemical Reactions

Change in intensive properties like color, odor, density

Release or absorption of energy

Signs of Chemical Reactions

Formation of a precipitate

Precipitate – *insoluble* solid formed from the reaction between 2 aqueous solutions

Chemical Reaction and Equation

Chemical reaction – a change that forms new substances

Reactants – starting substances

Products – new substances formed

Chemical Equation

- Shorthand way of describing chemical reactions

Example:

$$2H_2 + O_2 \rightarrow 2H_2O$$

(Reactants) (Product)

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

What are the reactants in the above reaction?

What are the products?

Types of Chemical Reactions

1. Combination or Synthesis

2 or more reactants → 1 product

Example: $H_2 + O_2 \rightarrow H_2O$

2. Decomposition

1 reactant → 2 or more products

Example: $H_2O_2 \rightarrow H_2O + O_2$

Types of Chemical Reactions

3. Single Displacement

2 reactants → 2 products

(active element and compound)

Example: $Mg + HCl \rightarrow MgCl_2 + H_2$

4. Double Displacement

2 reactants → 2 products

(2 aqueous solutions)

Example: $KI + Pb(NO_3)_2 \rightarrow KNO_3 + PbI_2$

Types of Chemical Reactions

5. Combustion

fuel + oxygen → water + carbon dioxide

Example: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

Types of Chemical Reactions

Classify each reaction:

1. $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$
2. $Na + H_2O \rightarrow NaOH + H_2$
3. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
4. $Mg + O_2 \rightarrow MgO$
5. $Na_2CO_3 \rightarrow Na_2O + CO_2$
6. $KOH + HCl \rightarrow KCl + H_2O$

Learning Objective:

- Write chemical equations from word equations.

Chemical Equations

Steps in writing chemical equations:

1. Identify the reactants and the products.

2. Write the formulae (or symbols) of the reactants before the arrow.

3. Write the formulae (or symbols) of the products after the arrow.

Chemical Equations

Example:

When magnesium (Mg) is heated, it reacts with oxygen (O_2) in air and burns to produce magnesium oxide (MgO).

Reactants: magnesium (Mg) and oxygen (O_2)

Products: magnesium oxide (MgO)

Chemical Equation: $Mg + O_2 \rightarrow MgO$

Write the chemical equation for this reaction:

Blue copper(II) sulfate solution reacts with iron to form iron(II) sulfate and copper.

Reactants:

Copper(II)sulfate and iron

Products:

Iron (II)sulfate and copper

Chemical Equation:

$CuSO_4 + Fe \rightarrow FeSO_4 + Cu$

Write the chemical equation for each reaction:

1. Silver oxide decomposes into silver and oxygen gas when heated.

2. Ethanol (C_2H_5OH) burns completely by reacting with oxygen in air. Carbon dioxide and water vapor are produced.

3. Aluminum bromide is produced when aluminum reacts with bromine.

Objective:

- Predict the products of common chemical reactions.

Predicting Products

$$Al + HCl \rightarrow \_\_\_\_\_$$

- What type of reaction will most likely occur between the 2 reactants?

- What are the products?

$$Al + HCl \rightarrow AlCl_3 + H_2$$

Predicting Products

$CuCl_2(aq) + Pb(NO_3)_2\ (aq) \rightarrow$ _____

- What type of reaction will most likely occur between the 2 reactants?

- What are the products?

$CuCl_2(aq) + Pb(NO_3)_2\ (aq) \rightarrow Cu(NO_3)_2\ (aq) + PbCl_2\ (s)$

Predicting Products

$$C_6H_{14} + O_2 \rightarrow \_\_\_\_\_$$

- What type of reaction will most likely occur between the 2 reactants?

- What are the products?

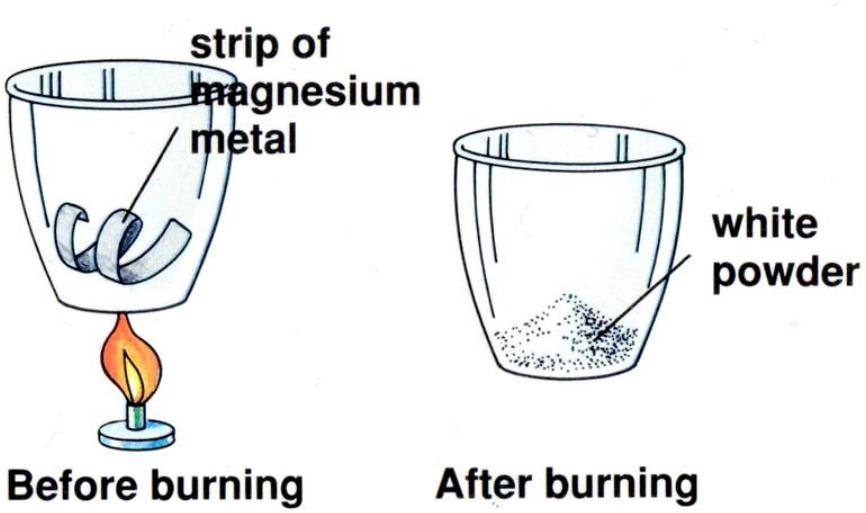
$$C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$$

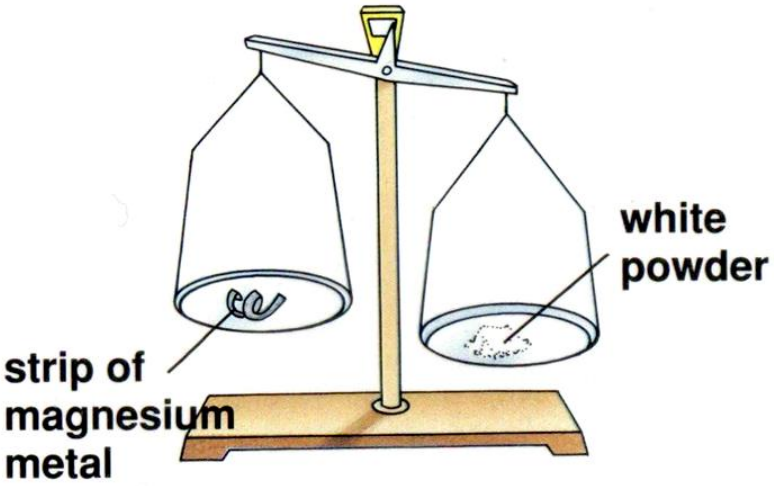
Predict the products of the reaction:

1. $NaOH(aq) + FeCl_3(aq) \rightarrow$
2. $Zn(s) + HNO_3(aq) \rightarrow$
3. $C_4H_{10}(g) + O_2(g) \rightarrow$
4. $N_2(g) + H_2(g) \rightarrow$
5. $KBr(aq) + Cl_2(_g) \rightarrow$

Law of Conservation of Mass

Burning Magnesium Metal in an Open Container

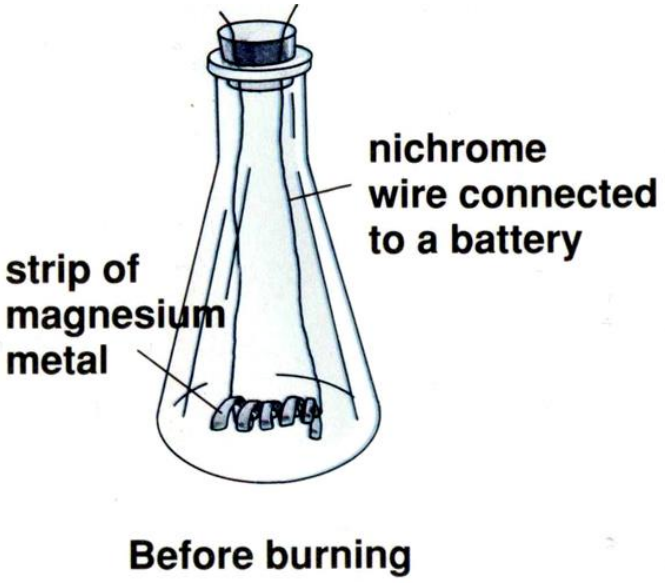


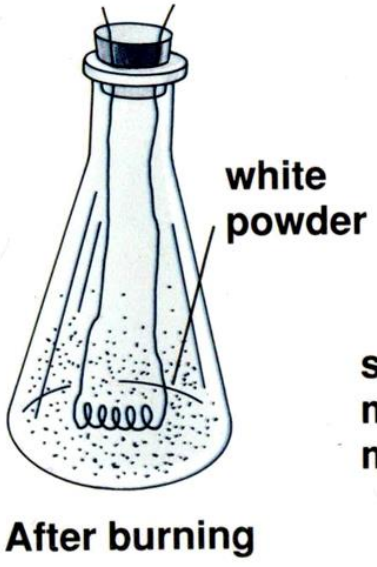


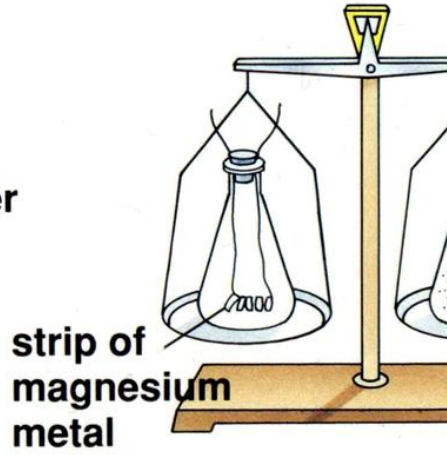
Dorin, Demmin, Gabel, Chemistry The Study of Matter , 3rd Edition, 1990, page 77

Law of Conservation of Mass

Burning Magnesium Metal in a Closed Container







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Law of Conservation of Mass

The total mass of reactants is equal to the total mass of the products.

Matter is neither created nor destroyed in a chemical reaction.

$$\begin{array}{ccccc} 2Mg & + & O_2 & \rightarrow & 2MgO \\ 48\ g & & 32\ g & & ? \end{array}$$

Balanced Chemical Equations

Balanced Equation – the number of atoms of each element is equal on both sides of the equation

How to balance equations:

1. Count the number of atoms of each element.
2. Use **coefficients** to make the number of atoms of each element equal.
3. DO NOT change any of the subscripts.

Balancing Chemical Equations

Example 1:

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

Reactants:

Mg – 1

H – 1 X 2 = 2

Cl – 1 X 2 = 2

Products:

Mg – 1

H – 2

Cl - 2

Balancing Chemical Equations

Example 2:

$$2\,Na + 2\,H_2O \rightarrow 2\,NaOH + H_2$$

Reactants:

Na – 1 X 2 = 2

H – 2 X 2 = 4

O – 1 X 2 = 2

Products:

Na – 1 X 2 = 2

H – 1 X 2 + 2 = 4

O – 1 X 2 = 2

Balancing Chemical Equations

Balance the following equations:

1. $Na + Cl_2 \rightarrow NaCl$
2. $Fe + O_2 \rightarrow Fe_2O_3$
3. $Zn + HCl \rightarrow ZnCl_2 + H_2$
4. $KNO_3 \rightarrow KNO_2 + O_2$

Closure: Write-Pair-Share

1. In your own words, describe how a chemical equation is balanced.

2. Share your answer with your group mates.

3. Make sure that everyone in the group has the correct answer to the question.

4. If your group is chosen and is able to give the correct answer, you earn 3 extra credit points.

Stoichiometry

-Stoichiometry is the calculation of the amount of reactants and products in a chemical reaction.

- Amount is usually expressed in number of moles, mass or volume (gases).

- Stoichiometric calculations are based on balanced equations.

Solving Stoichiometric Problems

1. Identify given and unknown.

2. Write possible conversion factors.

3. Set up equation using appropriate conversion factor(s).

$$\text{mole unknown} = \text{mole of known} \times \frac{\text{mole of unknown}}{\text{mole of known}}$$

Do the known units cancel?

4. Check answer. *Sig figs? Units?*

Mole-Mole Conversion

$2H_2 + O_2 \rightarrow 2H_2O$

2 moles **1 mole** **2 moles**

Coefficient – indicates number of moles

1.How many moles of oxygen are needed to produce 2 moles of water?

Answer: 1 mole of oxygen

2. How many moles of water can be produced from 4 moles of hydrogen?

Answer: 4 moles of water

3. How many moles of hydrogen is needed to react with 2 moles of oxygen?

Answer: 4 moles of hydrogen

Mole-Mole Conversion

$2H_2 + O_2 \rightarrow 2H_2O$

Ex.1 : How many moles of water can be produced from 3.50 moles of hydrogen?

Given: 3.50 mol H_2 Unknown: mol H_2O

Possible conversion factors: $\frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O}$ *Derived from balanced equation*

$\frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2}$

Equation:

$$\text{moles of } H_2O = 3.50 \text{ mol } H_2 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol of } H_2} = 3.50 \text{ mol } H_2O$$

Mole-Mole Conversion

$$2H_2 + O_2 \rightarrow 2H_2O$$

Ex. 2: How many moles of hydrogen are needed to react with 5.0 moles of oxygen?

Given: 5.0 mol O_2 Unknown: mol H_2

Possible conversion factors: $\frac{2\text{mol } H_2}{1 \text{ mol } O_2}$

$\frac{1 \text{ mol } O_2}{2\text{mol } H_2}$

Equation:

$$\text{moles of } H_2 = 5.0 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2}{1 \text{ mol of } O_2} = 10. \text{ moles } H_2$$

Practice Problems:

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

1. How many moles of magnesium are needed to produce 0.500 moles of magnesium chloride?

2. How many moles of hydrogen gas can be produced from 6 moles of magnesium?

Mole-Mass or Mass-Mole Conversion

Given

Molar mass

Mole ratio from balanced equation

$$\text{mol B} = \text{mass A} \times \frac{1\ \text{mol A}}{\text{mass A}} \times \frac{\text{mol B}}{\text{mol A}}$$

$$\text{mass B} = \text{mol A} \times \frac{\text{mol B}}{\text{mol A}} \times \frac{\text{mass B}}{1\ \text{mol B}}$$

Mole-Mass or Mass-Mole Conversion

$$2H_2 + O_2 \rightarrow 2H_2O$$

Ex.1 : What is the mass of oxygen that is needed to produce 4.0 moles of water?

Given: 4.0 mol H_2O Unknown: g of O_2

Possible conversion factors: $\frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O}$ $\frac{32g\ O_2}{1 \text{ mol } O_2}$

Equation:

$$\text{mass of } O_2 = 4.0 \text{ mol } H_2O \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} \times \frac{32g\ O_2}{1 \text{ mol } O_2} = 64 \text{ g } O_2$$

Practice Problems:

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

1. How many moles of magnesium are needed to form 47 grams of magnesium chloride?

2. How many grams of magnesium are needed to produce 4.5 moles of hydrogen?

Mass- Mass Conversion

$$2H_2 + O_2 \rightarrow 2H_2O$$

2(2g) = 4g 1(32g) = 32 g 2 (18g) = 36g

1. How many grams of hydrogen are needed to produce 36 g of water?

Answer: 4g hydrogen

2. How many grams of water can be produced from 32 g of oxygen?

Answer: 36 g water

3. What is the mass of oxygen that is needed to react with 8 g of hydrogen?

Answer: 2(32g) = 64 g

Mass -Mass Conversion

$$\text{mass B} = \text{mass of A} \times \frac{\text{1 mole A}}{\text{mass A}} \times \frac{\text{mole B}}{\text{mole A}} \times \frac{\text{mass B}}{\text{1 mole B}}$$

$$2H_2 + O_2 \rightarrow 2H_2O$$

Ex.1 : What is the mass of oxygen that is needed to produce 18 g of water?

Given: 18 g of H_2O Unknown: g of O_2

Possible conversion factors: $\frac{1 \text{ mol } H_2O}{18 \text{ g } H_2O}$ $\frac{1 \text{ mol } O_2}{2 \text{mol } H_2O}$ $\frac{32 \text{g } O_2}{1 \text{ mol } O_2}$

Equation:

$$\text{mass of } O_2 = 18 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18 \text{ g } H_2O} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} \times \frac{32 \text{ g } O_2}{1 \text{ mol } O_2} = 16 \text{ g } O_2$$

Practice Problems:

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

1. How many grams of magnesium are needed to produce 6g of hydrogen?

2. How many grams of magnesium chloride can be produced from 54 g magnesium?

Limiting and Excess Reactants

Limiting Reactant – completely used up; limits the amount of product

Excess Reactant – not completely used up, “left over”

Limiting and Excess Reactants

$$2H_2 + O_2 \rightarrow 2H_2O$$

Ex.1: 6.0 g of H_2 and 60.g of O_2 are made to react.

(a) Is there a reactant present in excess? If there is, how many grams of this reactant is left unreacted?

(b) How many grams of water are produced from the reaction?

$$\text{mass of } O_2 = 6.0 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2 \text{ g } H_2} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2} \times \frac{32 \text{g } O_2}{1 \text{ mol } O_2} = 48 \text{ g } O_2$$

Only 48 g of O_2 is needed to completely react, so it is an excess reactant. 12 g of O_2 is left over.

Limiting and Excess Reactants

$$2H_2 + O_2 \rightarrow 2H_2O$$

Ex.1: 6.0 g of H_2 and 60.g of O_2 are made to react.

(a) Is there a reactant present in excess? If there is, how many grams is left unreacted?

(b) How many grams of water are produced from the reaction?

H_2 is the limiting reactant; it determines the amount of water produced.

$$\text{mass of } H_2O = 6.0 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2 \text{ g } H_2} \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} \times \frac{18 \text{ g } H_2O}{1 \text{ mol } H_2O} = 54\text{g } H_2O$$

Practice Problems:

$$2Al + 3Br_2 \rightarrow 2AlBr_3$$

20 g aluminum and 100.0 g bromine were made to react.

1. What is the limiting reactant in the reaction?

2. How much of the excess reactant is left over after the reaction?

3. How many grams of aluminum bromide is produced from the reaction?

Balancing Chemical Equations

- Follows the Law of Conservation of Mass.
- Steps to balance equations:
- 1. Write the unbalanced equation.
- 2. Count atoms of each element in reactants and products.
- 3. Adjust coefficients to balance atoms for each element.
- 4. Ensure all coefficients are in the simplest whole-number ratio.
- 5. Check the final balanced equation.

Example: Balancing Equations

- Unbalanced: $H_2 + O_2 \rightarrow H_2O$
- Balanced: $2H_2 + O_2 \rightarrow 2H_2O$
- All atoms are now equal on both sides.

Stoichiometric Quantities

- Stoichiometry: Quantitative relationship between reactants and products.
- Mole Ratio: Ratio of moles of one substance to another in a balanced equation.
- Steps for Stoichiometric Calculations:
- 1. Convert given mass or volume to moles.
- 2. Use mole ratio from balanced equation to relate reactants to products.
- 3. Convert moles back to mass or volume if required.

Example: Stoichiometry Calculation

- Given: 4.00 g of hydrogen reacts with oxygen to form water.

- Find: Mass of water produced.
- 1. Convert mass of H_2 to moles: 4.00 g / 2.016 g/mol = 1.98 moles H_2.
- 2. Use mole ratio: $2H_2 + O_2 \rightarrow 2H_2O$.
- 3. 1.98 moles H_2 will produce 1.98 moles H_2O.
- 4. Convert to grams: 1.98 × 18.016 g/mol = 35.7 g H_2O.

Example: Stoichiometry Calculation

- Given: 10.0 g of propane (C_3H_8) reacts with oxygen.
- Find: Mass of CO_2 produced.
- 1. Convert mass of C_3H_8 to moles: 10.0 g / 44.1 g/mol = 0.227 moles C_3H_8.
- 2. Use mole ratio: 1 C_3H_8 → 3 CO_2.
- 3. 0.227 moles C_3H_8 will produce 0.227 × 3 = 0.681 moles CO_2.
- 4. Convert to grams: 0.681 × 44.01 g/mol = 29.96 g CO_2.

Limiting and Excess Reactants

- Limiting Reactant: The reactant that is completely consumed first.
- Excess Reactant: The reactant that remains after the reaction is complete.
- Steps to Identify the Limiting Reactant:
- 1. Convert given amounts of reactants to moles.
- 2. Use mole ratio to find the required amount of one reactant for the given amount of the other.
- 3. The reactant that produces the least amount of product is the limiting reactant.

Example: Limiting Reactant

- Given: 5.00 g H_2 and 10.0 g O_2.
- Balanced equation: $2H_2 + O_2 \rightarrow 2H_2O$.
- Convert reactants to moles:
- - H_2: 5.00 g / 2.016 g/mol = 2.48 moles H_2.
- - O_2: 10.0 g / 32.00 g/mol = 0.3125 moles O_2.
- O_2 is the limiting reactant since it produces the least amount of H_2O.

Example: Limiting Reactant

- Given: 10.0 g of aluminum reacts with 15.0 g of chlorine gas to form $AlCl_3$.
- Balanced equation: $2Al + 3Cl_2 \rightarrow 2AlCl_3$.
- Convert reactants to moles:
- - Al: 10.0 g / 26.98 g/mol = 0.371 moles.
- - Cl_2: 15.0 g / 70.90 g/mol = 0.211 moles.
- Use mole ratio to determine limiting reactant:
- - Required Cl_2 for 0.371 moles Al: 0.371 × (3/2) = 0.557 moles.
- - Since only 0.211 moles Cl_2 is available, Cl_2 is the limiting reactant.

Excess Reactant Calculation

- Remaining reactant = Initial amount - Required amount.
- Since O_2 is the limiting reactant, H_2 is in excess.
- To find excess H_2:
- 1. Required H_2 for 0.3125 moles O_2: (0.3125 × 2) = 0.625 moles H_2.
- 2. Excess H_2 = 2.48 - 0.625 = 1.86 moles.
- 3. Convert to grams: 1.86 × 2.016 g/mol = 3.75 g H_2 remains unreacted.

Example: Balancing Chemical Equations

- Unbalanced: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
- Balanced: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
- Explanation: Carbon, hydrogen, and oxygen atoms are balanced on both sides.

Summary

- Chemical equations must be balanced.
- Stoichiometry helps determine reactant-product relationships.
- Limiting reactant determines the maximum product formation.
- Excess reactant remains unreacted.

Theoretical and Percent Yield

Theoretical Yield – amount of product formed when all of the reactants are completely used up

Actual Yield – amount of product actually formed in a reaction

Usually: Actual Yield < Theoretical Yield

Percent Yield – indicates how well a reaction comes to completion

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

Sample Problem 1

$2H_2$	+	O_2	→	$2H_2O$
2(2g) = 4g		1(32g) = 32 g		2 (18g) = 36g

1. How many grams of water can be produced from 32 g of oxygen?

Answer: 36 g water (theoretical yield)

2. If only 27 g of water is actually produced from 32 g of oxygen, what is the % yield of the reaction?

Answer: % yield = $\frac{27}{36} \times 100 = 75\%$

Sample Problem 2

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

1. How many grams of hydrogen are formed from 24 g magnesium?

Answer: 2 g

2. If only 1g of hydrogen is actually produced from 24 g magnesium, what is the % yield of the reaction?

Answer: $\%\ \text{yield} = \frac{1\text{g}}{2\text{g}} \times 100 = 50\%$

Check for Understanding

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

What is the % yield of the above reaction if only 45 g of oxygen is produced from 122 g of potassium chlorate?

Questions & Discussion

- Practice balancing equations.
- Solve a stoichiometry problem.
- Identify limiting reactants in a reaction scenario.