Unit 1
CHEMICAL REACTIONS AND STOICHIOMETRY

1.1 Introduction

Change is an inevitable part of nature. We encounter various changes in our everyday lives, such as the growth of a plant from a seed, the burning of wood, the rusting of iron, the rotting of food, and the evaporation of liquids. Scientists classify these changes into two main categories: physical changes and chemical changes.

Physical Change:
A physical change does not involve the formation of new substances. The chemical composition of the material remains the same. Examples of physical changes include:

- **Evaporation** of water (liquid to gas)
- **Powdering** of sugar (solid to fine particles)
- **Melting** of ice (solid to liquid)

Chemical Change:
A chemical change results in the formation of one or more new substances with a new chemical composition. Chemical changes involve chemical reactions, where reactants are transformed into products with different properties. Examples of chemical changes are:

- **Turning milk into curd** (a biochemical process involving bacteria)
- **Photosynthesis** in plants (conversion of carbon dioxide and water into glucose and oxygen)
- **Rotting of eggs** (decomposition by bacteria)

Characteristics of Chemical Reactions:

1. **Formation of New Substances:** New substances with different properties are formed.
2. **Production of Heat or Light:** Some reactions release energy in the form of heat or light.
3. **Change in Color:** The color of the substances may change.
4. **Change in Temperature:** The reaction may result in a temperature change.

A chemical reaction can be represented by a **chemical equation** which uses symbols to show what happens during the reaction. For example, burning sulfur with oxygen forms sulfur dioxide, which has distinct properties from sulfur and oxygen.

1.2 Chemical Equations

A chemical equation is a shorthand representation of a chemical reaction using chemical symbols and formulas. The substances that react are called reactants, and the new substances formed are called products.

Writing Chemical Equations:

1. **Word Equation:** Describes the reaction in words.
   - Example: Hydrogen + Nitrogen → Ammonia
2. **Chemical Equation:** Uses symbols and formulas to represent the reaction.
   - Example: $H_2 + N_2 \rightarrow NH_3$
3. **Balanced Chemical Equation:** The number of atoms for each element is the same on both sides of the equation.
   - Example: $3H_2 + N_2 \rightarrow 2NH_3$

Indicators of Physical States:

- **(s)** - Solid
- **(l)** - Liquid
- **(g)** - Gas
- **(aq)** - Aqueous solution (dissolved in water)

Example 1.1:
Reaction Between Calcium Carbonate and Sulfuric Acid:

1. **Word Equation:** Calcium carbonate + Sulfuric acid → Calcium sulfate + Water + Carbon dioxide
2. **Chemical Equation:** $CaCO_3(s) + H_2SO_4(l) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$
3. **Balanced Equation:** The equation is balanced as written.

Example 1.2:
Reaction Between Sodium Chloride and Silver Nitrate:

1. **Word Equation:** Silver nitrate + Sodium chloride → Silver chloride + Sodium nitrate
2. **Chemical Equation:** $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
3. **Balanced Equation:** The equation is balanced as written.

Conditions for a Balanced Chemical Equation:

1. **True and Possible Reaction:** The equation must represent a real chemical reaction.
2. **Correct Symbols and Formulas:** Chemical symbols and formulas must be accurate.
3. **Diatomic Molecules:** Some elements (e.g., H_2, O_2) exist as diatomic molecules in equations.
4. **Balanced Equation:** The number of atoms for each element must be equal on both sides.

1.2.1 Balancing Chemical Equations

Law of Conservation of Mass: Atoms are neither created nor destroyed in a chemical reaction. The number of atoms of each element must be the same on both sides of the equation.

Balancing Chemical Equations by Inspection Method:

1. **Write the Word Equation:** Describe the reaction in words.
2. **Write the Unbalanced Chemical Equation:** Represent the reactants and products with symbols.
3. **Count Atoms:** Tabulate the number of each type of atom on both sides.
4. **Balance Atoms:** Adjust coefficients to ensure the number of atoms is equal on both sides.

Example of Balancing by Inspection Method:

- **Unbalanced Equation:** $Fe + H_2O \rightarrow Fe_3O_4 + H_2$
- **Balanced Equation:** $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

Balancing Chemical Equations by the LCM Method:

1. **Write the Word Equation:** Describe the reaction in words.
2. **Write the Chemical Symbols:** Represent reactants and products with formulas.
3. **Determine Total Valency:** Calculate the valency of each element in reactants and products.
4. **Find LCM:** Use the Least Common Multiple (LCM) to balance the equation.
5. **Substitute Coefficients:** Adjust coefficients to balance the atoms.

Example of LCM Method:

- **Reaction:** Aluminum + Oxygen → Aluminum Oxide

- **Unbalanced Equation:** $Al + O_2 \rightarrow Al_2O_3$
- **Balanced Equation:** $4Al + 3O_2 \rightarrow 2Al_2O_3$

Balancing Chemical Equations Using Algebraic Method:

1. **Write the Unbalanced Equation:** Include symbols for reactants and products.
2. **Assign Variables:** Use algebraic variables for coefficients.
3. **Set Up Equations:** Balance the atoms by setting up equations for each element.
4. **Solve Equations:** Determine the values of the variables and substitute them back into the equation.

Example of Algebraic Method:

- **Unbalanced Equation:** $N_2 + H_2 \rightarrow NH_3$
- **Balanced Equation:** $1N_2 + 3H_2 \rightarrow 2NH_3$

1.3 Types of Chemical Reactions

Chemical reactions can be classified into four basic types:

1. **Combination Reactions:** Two or more substances combine to form a single product.
   - **Example:** $2Na + Cl_2 \rightarrow 2NaCl$
2. **Decomposition Reactions:** A single compound breaks down into two or more products.
   - **Example:** $2H_2O \rightarrow 2H_2 + O_2$
3. **Single Displacement Reactions:** One element replaces another in a compound.
   - **Example:** $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
4. **Double Displacement Reactions:** The ions of two compounds exchange places in an aqueous solution.
   - **Example:** $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

Understanding these reaction types helps predict the products of unknown reactions and classify them accordingly.

1.4 Oxidation and Reduction Reactions

Activity 1.3

Objective: To understand the concepts of oxidizing and reducing agents through everyday examples.

Scenario: When drying dishes with a towel, the towel absorbs water from the dishes.

Discussion Points:

- **Drying Agent vs. Wetting Agent:** The towel is the drying agent (which absorbs water) and the dishes are the wetting agent (which loses water).
- **Oxidizing and Reducing Agents Analogy:** Just like the towel absorbs water (loses water to the towel), in a redox reaction, one substance loses electrons (is oxidized) and another gains electrons (is reduced).

Conclusion: The towel (drying agent) can be related to a reducing agent because it gains water (or electrons), while the dishes (wetting agent) can be related to an oxidizing agent because it loses water (or electrons).

Oxidation-Reduction

Oxidation is the loss of one or more electrons by an atom. **Reduction** is the gain of one or more electrons. These processes occur simultaneously in a reaction known as a redox reaction.

Example: When zinc reacts with copper(II) sulfate: $Zn+CuSO_4 \rightarrow ZnSO_4+Cu$ In this reaction:

- **Zinc (Zn)** loses electrons and is oxidized.
- **Copper (Cu)** gains electrons and is reduced.

Chemical Equations: $Zn+Cu^{2+} \rightarrow Zn^{2+}+Cu$

Oxidation and reduction can also be defined by changes in oxidation numbers.

Example: $Cu_0+2Ag^+ \rightarrow Cu^{2+}+2Ag_0$

- Copper’s oxidation number increases from 0 to +2 (oxidized).
- Silver’s oxidation number decreases from +1 to 0 (reduced).

1.4.1 Oxidation Number or Oxidation State

Definition: The oxidation number (or state) indicates the number of electrons an atom has gained or lost when combined with other atoms.

Rules for Assigning Oxidation Numbers:

1. **Uncombined Elements:** Oxidation number is 0.
    - Examples: Be = 0, O in O_2 = 0.
2. **Monatomic Ions:** Oxidation number equals the ion's charge.
    - Examples: Na^+ = +1, S^{2-} = -2.
3. **Oxygen:** Usually -2, except in peroxides (-1), superoxides (-1/2), and with fluorine (+2).
    - Examples: Na_2O_2 (peroxide) = -1.
4. **Hydrogen:** Usually +1, except in metal hydrides where it is -1.
    - Examples: NaH (metal hydride) = -1.
5. **Sum in Neutral Compounds:** Total oxidation number is 0.
    - Example: H_2SO_4 : +1 + 6 - 8 = 0.
6. **Sum in Polyatomic Ions:** Equals the ion's charge.
    - Example: SO_4^{2-} : +6 + (-8) = -2.
7. **Group IA Elements:** Always +1.
8. **Group IIA Elements:** Always +2.

Examples:

- **HNO_3:** H = +1, N = +5, O = -2.
- **CrO_4^{2-} :** Cr = +6, O = -2.
- **MnO_4^- :** Mn = +7, O = -2.
- **$Ca(H_2PO_4)_2$:** Ca = +2, P = +5, O = -2.

1.4.2 Oxidizing and Reducing Agents

Oxidizing Agent: Gains electrons (causes oxidation).

- **Characteristics:**
    1. Gains electrons.
    2. Causes oxidation.
    3. Undergoes reduction.
    4. Becomes more negative.

Reducing Agent: Loses electrons (causes reduction).

- **Characteristics:**
    1. Loses electrons.
    2. Causes reduction.
    3. Undergoes oxidation.
    4. Becomes more positive.

Tests for Agents:

- **Oxidizing Agent:** Changes color with substances easily oxidized.
- **Reducing Agent:** Changes color with substances easily reduced.

Examples:

- **Reducing Agent:** Potassium iodide (KI) turns blue-black with iodine.
- **Oxidizing Agent:** Hydrogen sulfide (H_2S) forms yellow precipitate with oxidizing agents.

Factors Affecting Agent Strength:

- **Electronegativity:** High electronegativity elements are good oxidizing agents.
- **Oxidation States:** Higher oxidation states indicate oxidizing agents; lower oxidation states indicate reducing agents.

1.4.3 Analyzing Redox Reactions

To Identify Agents in Reactions:

1. Determine which species loses electrons (oxidized) and which gains electrons (reduced).
2. An oxidizing agent undergoes reduction; a reducing agent undergoes oxidation.
3. Every redox reaction must have both an oxidizing agent and a reducing agent to ensure the transfer of electrons.

1.4.4 Balancing Redox Reactions: Oxidation-Number-Change Method

Steps to Balance Redox Reactions:

1. **Assign Oxidation Numbers:** Determine changes in oxidation states of reactants and products.
2. **Identify Changes:** Recognize which elements are oxidized and which are reduced.
3. **Connect Changes:** Use lines to show changes in oxidation numbers.
4. **Balance Changes:** Ensure total electron loss equals total electron gain by adjusting coefficients.

Example: $Fe_2O_3 + CO \rightarrow Fe + CO_2$

- **Oxidation Number Changes:**
  - Fe: +3 to 0 (reduction).
  - C: +2 to +4 (oxidation).

- **Balanced Equation:** $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1.5 Chemical Quantities

1.5.1 Molecular Mass (MM) and Formula Mass (FM)

Molecular Mass (MM): Sum of the atomic masses of all atoms in a molecule.

Formula Mass (FM): Sum of atomic masses in a formula unit.

Example: The molecular mass of H2O is used to calculate the percentage composition of hydrogen and oxygen.

Percent Composition by Mass: Percent Composition=$\left(\frac{n \times Molar\ Mass\ of\ Element}{\text{Molar Mass of Compound}}\right) \times 100\%$

Example Calculation: For H_2O_2:

- % H = $\frac{2 \times 1.008\ g}{34.02\ \text{g}} \times 100\% = 5.926\%$
- % O = $\frac{2 \times 16.00\ g}{34.02\ \text{g}} \times 100\% = 94.074\%$

1.5.2 Empirical and Molecular Formulas

Empirical Formula: Shows the simplest whole-number ratio of elements in a compound.

Molecular Formula: Shows the actual number of atoms of each element in a molecule.

Steps to Determine Formulas:

1. Calculate the empirical formula from percentage composition or mass.
2. Determine the molecular formula using empirical formula and molecular mass.

Example Calculation:

- **Given:**
  - Molecular mass of a compound is 180 g/mol.
  - Empirical formula is CH_2O.
- **Calculate:**

- Empirical formula mass: C = 12.01, H = 2.016, O = 16.00.
- Empirical formula mass = 30.026 g/mol.
- Molecular formula = $\frac{180\ g/mol}{30.026\ \mathrm{g/mol}}$ = 6. Therefore, molecular formula is $C_6H_{12}O_6$.

1.5.3 The Mole Concept

Mole: A mole is a quantity that contains 6.022×10^{23} particles (Avogadro's number).

Molar Mass: The mass of one mole of a substance, expressed in grams per mole (g/mol).

Converting Moles to Mass: Mass=Number of Moles×Molar Mass

Converting Mass to Moles: Number of Moles=$\frac{Mass}{\mathrm{Molar\ Mass}}$

Example:

- To find the mass of 2 moles of NaCl: Molar Mass of NaCl=58.44 g/mol
  Mass=2 moles×58.44 g/mol=116.88 g

1.6 Stoichiometry

Stoichiometry is the branch of chemistry that studies the quantitative relationships between the substances involved in chemical reactions. It involves calculations of the amounts of reactants and products in chemical reactions based on the balanced chemical equations. Stoichiometric calculations are guided by two main principles:

A. Composition and Conservation of Mass

1. **Substance Composition**: Each substance in a chemical reaction is represented by a definite formula in the balanced equation.
2. **Law of Conservation of Mass**: The mass of reactants in a chemical reaction equals the mass of the products.

1.6.1 Molar Ratios in Balanced Chemical Equations

In stoichiometry, we use balanced chemical equations to determine the relationships between the amounts of reactants and products. For example, the reaction of hydrogen and nitrogen to produce ammonia can be represented as:

$N_2 + 3H_2 \rightarrow 2NH_3$

This equation indicates:

- **Molecular Interpretation**: 1 molecule of N_2 reacts with 3 molecules of H_2 to produce 2 molecules of NH_3.
- **Molar Interpretation**: 1 mole of N_2 reacts with 3 moles of H_2 to produce 2 moles of NH_3.
- **Mass Interpretation**: 28.0 g of N_2 reacts with 6.06 g of H_2 to produce 34.0 g of NH_3.

1.6.2 Mass-Mass Relationships

A. The Mass-Ratio Method

1. **Write the Balanced Equation**: Ensure the chemical equation is balanced.
2. **Place the Given Mass**: Write the given mass above the corresponding substance in the equation.
3. **Calculate Molar Masses**: Determine the molar masses and write them below the formulas.
4. **Set Up Proportions**: Use the molar masses and coefficients to set up a proportion.
5. **Solve for Unknown**: Calculate the unknown mass using the proportion.

B. The Mole-Ratio Method

1. **Write the Balanced Equation**: Ensure the equation is balanced.
2. **Convert Mass to Moles**: Convert the given mass of one substance to moles.
3. **Set Up the Mole Ratio**: Use the mole ratios from the balanced equation.
4. **Solve for Unknown**: Find the moles or mass of the required substance and convert if necessary.

Example: Mass-Mass Problem

For the reaction: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Find grams of CO required to produce 28 g of Fe:

1. **Calculate Moles of Fe**: Moles of Fe$=\frac{28\ g}{56\ \text{g/mol}}$ =0.5 mol
2. **Use Mole Ratio**: $\frac{0.5\ mol\ Fe}{2\ \text{mol Fe}} = \frac{x}{3\ \text{mol CO}}$
   - x=0.75 mol CO
3. **Convert Moles to Grams**: Mass of CO=0.75 mol×28 g/mol=21 g

1.6.3 Volume-Volume Relationships

At Standard Temperature and Pressure (STP), 1 mole of any gas occupies 22.4 liters. Avogadro's Law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

In volume-volume problems, you calculate the volume of one gas given the volume of another gas.

1.6.4 Mass-Volume Relationships

In these problems, you either find the volume of a substance from its mass or vice versa. The steps involve:

1. **Converting Mass to Moles**: Use molar mass.
2. **Using Stoichiometric Ratios**: Apply the balanced chemical equation.
3. **Converting Moles to Volume**: Use molar volume at STP if needed.

1.6.5 Limiting and Excess Reactants

In a reaction, the limiting reactant is the one that runs out first and limits the amount of product formed. The excess reactant remains after the reaction is complete.

Example: If you have 7 slices of bread and 2 slices of cheese to make sandwiches:

- You can make 2 sandwiches.
- 3 slices of bread are left over.
- Cheese is the limiting reactant; bread is in excess.

Example Calculation: For the reaction of Mg with HCl:

- Determine the limiting reactant by calculating the amount of product produced from each reactant.
- Use the limiting reactant to find the yield of the product.

1.6.6 Theoretical, Actual, and Percentage Yields

- **Theoretical Yield**: The amount of product expected based on the stoichiometric calculations.
- **Actual Yield**: The amount of product actually obtained from the reaction.
- **Percentage Yield**: The ratio of the actual yield to the theoretical yield, multiplied by 100.

Percentage Yield=$\left(\frac{\textit{Actual Yield}}{\text{Theoretical Yield}}\right)\times 100\%$

Conclusion: Not all chemical reactions proceed with 100% yield due to side reactions, incomplete reactions, or losses during product collection. Theoretical yield is calculated based on stoichiometry, while actual yield is measured experimentally. The percentage yield reflects the efficiency of the reaction.