Chapter 1: Chemical Reactions and Equations

Introduction to Chemical Reactions

In our everyday lives, we observe many changes that indicate a chemical reaction has taken place. For example:

- Milk turns sour when left at room temperature during summer due to microbial activity producing lactic acid.
- Iron utensils or nails left in a humid atmosphere start forming a reddish-brown coating, known as rust. This is the result of iron reacting with oxygen and moisture.
- Grapes ferment due to yeast converting sugars into alcohol and carbon dioxide.
- Cooking food changes raw ingredients into cooked food by breaking down complex molecules and combining new ones.
- In our body, complex food molecules like carbohydrates, proteins, and fats are broken down into simpler absorbable forms during digestion.
- Respiration is a chemical reaction in which glucose reacts with oxygen to release energy needed for body functions.

These examples show **chemical changes**, where the original substances transform into new substances with different properties. Such changes are known as **chemical reactions**.

How to Identify a Chemical Reaction Has Occurred

Certain signs indicate that a chemical reaction has taken place:

1. **Change in State**: A solid forming from a solution or gas forming from a liquid.
    - Example: Zinc reacting with hydrochloric acid produces hydrogen gas (gas evolution).
2. **Change in Colour**: A substance changes its colour during the reaction.

    - Example: Iron nail in copper sulfate solution causes the blue colour to fade.
3. **Evolution of Gas**: Bubbles form, showing a gas is released.
    - Example: Reaction of marble chips with hydrochloric acid releases carbon dioxide.
4. **Change in Temperature**: Heat is either released or absorbed.
    - Example: Calcium oxide reacts with water producing heat (exothermic reaction).

These are observable clues that help in confirming whether a chemical reaction has occurred.

Activity 1.1: Burning Magnesium Ribbon

When a magnesium ribbon is cleaned and burned in air:

- It burns with a bright white flame.
- A white powder is formed and collected in a watch-glass.
- This powder is magnesium oxide (MgO).

This activity is used to demonstrate a simple chemical reaction where magnesium (a metal) reacts with oxygen (from the air) to form a compound (magnesium oxide). This is a **combination reaction** and also an **exothermic reaction** as it releases heat and light.

Chemical Equations

A **chemical equation** represents a chemical reaction using symbols and formulas.

Word Equation

Magnesium + Oxygen → Magnesium Oxide

This represents the names of the substances involved.

Skeletal Chemical Equation

$Mg + O_2 \rightarrow MgO$

This is the unbalanced form of a chemical equation. It is called skeletal because it only shows the basic form using symbols and formulas without considering the Law of Conservation of Mass.

1.1.1 Writing a Chemical Equation

Chemical equations are a shorthand representation of chemical reactions. They show:

- **Reactants**: Substances that react and are written on the left-hand side.
- **Products**: Substances that are formed and are written on the right-hand side.
- The arrow (→) points from reactants to products.

Chemical equations are more informative and concise than word equations.

1.1.2 Balanced Chemical Equations

Balancing a Chemical Equation Step-by-Step:

Example: $Fe + H_2O \rightarrow Fe_3O_4 + H_2$

Step I: List the number of atoms on both sides:

- Fe: 1 (LHS), 3 (RHS)
- H: 2 (LHS), 2 (RHS)
- O: 1 (LHS), 4 (RHS)

Step II: Balance oxygen by adjusting H_2O molecules to 4:
$Fe + 4H_2O \rightarrow Fe_3O_4 + H_2$

Step III: Balance hydrogen: $4H_2O$ contains 8 hydrogen atoms. So we need $4H_2$ on RHS.
$Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

Step IV: Balance iron: Fe_3O_4 has 3 iron atoms, so write 3Fe on LHS:
$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

This is now a **balanced chemical equation**, meaning it respects the Law of Conservation of Mass.

Physical States in Equations

To make chemical equations more informative, the physical states of the reactants and products are included:

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

Example:
$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$

Here, water is in the form of steam (gas) and iron is solid.

Conditions of Reaction

Sometimes, chemical reactions need special conditions like temperature, pressure, or the presence of a catalyst. These are indicated above or below the arrow in a chemical equation.

Example:
$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$
(With high pressure and catalyst)

Photosynthesis is another example:
$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
(In the presence of sunlight and chlorophyll)

1.2 Types of Chemical Reactions

Chemical reactions can be classified into different types based on the nature of the transformation:

1.2.1 Combination Reaction

In a combination reaction, two or more substances combine to form a single new substance.

Example 1:
$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + Heat$
(Quick lime + Water → Slaked lime)

This reaction releases heat, making it exothermic. It is also used in whitewashing.

Example 2:
$C(s) + O_2(g) \rightarrow CO_2(g)$
(Carbon burns in oxygen to form carbon dioxide)

Example 3:
$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
(Hydrogen and oxygen combine to form water)

1.2.2 Decomposition Reaction

In this reaction, a single substance breaks down into two or more simpler substances.

Decomposition reactions require energy, and can be:

- **Thermal decomposition** (using heat)
- **Photolytic decomposition** (using light)
- **Electrolytic decomposition** (using electricity)

Example 1 (Thermal):
$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
(Calcium carbonate → Calcium oxide + Carbon dioxide)

Example 2 (Photolytic):
$2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$
(Silver chloride decomposes in sunlight)

Example 3 (Electrolytic):
$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
(Electrolysis of water)

1.2.3 Displacement Reaction

In displacement reactions, a more reactive element displaces a less reactive one from its compound.

Example:
$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
(Iron displaces copper from copper sulfate solution)

Other examples:

- $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
- $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$

1.2.4 Double Displacement Reaction

In this reaction, two compounds exchange ions to form new compounds. One of the products is often a precipitate.

Example:
$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

$BaSO_4$ is an insoluble white precipitate.

1.2.5 Oxidation and Reduction (Redox Reactions)

- **Oxidation**: Gain of oxygen or loss of hydrogen.
- **Reduction**: Loss of oxygen or gain of hydrogen.

These reactions happen together. One substance gets oxidised while the other gets reduced. Such reactions are called **Redox reactions**.

Example:
$2Cu + O_2 \rightarrow 2CuO$ (Oxidation of copper)

$CuO + H_2 \rightarrow Cu + H_2O$ (Reduction of copper oxide)

Other Examples:

- $ZnO + C \rightarrow Zn + CO$ (ZnO is reduced; C is oxidised)
- $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

1.3 Effects of Oxidation in Everyday Life

Chemical reactions involving oxidation not only occur in laboratories but also impact our daily lives. Two common examples of oxidation-related processes are **corrosion** and **rancidity**.

1.3.1 Corrosion

Corrosion is the process by which metals get deteriorated when they react with substances in their environment such as oxygen, water, acids, etc. It is essentially a slow oxidation process of metals.

Example: Rusting of Iron
When iron reacts with moisture and oxygen in the air, it forms a reddish-brown flaky substance called **rust** (chemically $Fe_2O_3 \cdot xH_2O$).

Reaction: (which eventually dehydrates to form rust)

Other Examples of Corrosion:

- Silver reacts with sulfur in the air to form a black coating of silver sulfide (Ag_2S).
- Copper develops a green coating of basic copper carbonate due to prolonged exposure to moist air.

Consequences of Corrosion:

- Weakening of structures like bridges, railings, pipelines, and vehicles.
- Loss of appearance and mechanical strength.
- Huge financial losses due to maintenance and replacements.

Prevention of Corrosion:

- Painting or coating the metal surface.
- Applying oil or grease.
- Galvanization (coating with zinc).
- Electroplating.
- Using anti-rust solutions.

- Storing in dry environments.

1.3.2 Rancidity

Rancidity refers to the condition in which fats and oils in food items get oxidised, resulting in a bad taste and unpleasant smell.

This is also a result of **oxidation** of fats and oils when exposed to oxygen in the air.

Example: Food containing oil or ghee (like chips, fried snacks) becomes inedible after being exposed to air for a few days.

Types of Rancidity:

- **Oxidative rancidity**: Caused by the oxidation of unsaturated fatty acids.
- **Hydrolytic rancidity**: Caused by the hydrolysis of fats in the presence of water.

Prevention of Rancidity:

- Adding **antioxidants** (substances that prevent oxidation), e.g., BHA and BHT.
- Storing food in **airtight containers** to prevent exposure to oxygen.
- Flushing packets with **nitrogen gas** (as done in chip packets) to remove oxygen.
- **Refrigerating** or storing food in a cool and dark place to slow down oxidation.

Summary of Key Concepts

- Chemical reactions lead to the formation of new substances with new properties.
- Word and chemical equations represent these reactions, and equations must be balanced.
- There are different types of chemical reactions: Combination, Decomposition, Displacement, Double Displacement, and Redox.

- Some reactions are **exothermic** (release heat) and others are **endothermic** (absorb heat).
- Redox reactions involve simultaneous oxidation and reduction.
- **Corrosion** and **rancidity** are oxidation reactions that affect metals and food respectively.