Unit 4

ENERGY CHANGES AND ELECTROCHEMISTRY

4.1 Introduction to Energy Changes in Chemical Reactions

Chemical reactions often involve a change in energy, meaning the energy of the system before and after the reaction is different. For example, when carbon (found in coke) burns in oxygen, it forms carbon dioxide and releases energy as heat and light:

$C(s) + O_2(g) \rightarrow CO_2(g) + \text{energy (heat and light)}$

Similarly, the combustion of methane (CH_4) in oxygen produces carbon dioxide, water, and energy:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) + \text{energy (heat and light)}$

To study these energy changes (ΔE), scientists define two parts of the universe: the **system** (the part being studied, such as the reaction mixture) and the **surroundings** (everything else). The energy within the system is called **internal energy (E)**, which is the sum of its potential and kinetic energy:

$E_{system} = \text{Potential Energy} + \text{Kinetic Energy}$

When reactants convert to products in a chemical reaction, the system's internal energy changes:

$\Delta E = E_{final} - E_{initial} = E_{products} - E_{reactants}$

Key Concepts:

- **System**: The part of the universe being studied (e.g., reaction mixture).
- **Surroundings**: Everything else in the universe.
- **Internal Energy (E)**: Total energy within the system (sum of potential and kinetic energy).
- **ΔE (Change in Energy)**: Difference in internal energy before and after the reaction.

Energy Transfer Between System and Surroundings

Energy is transferred between the system and its surroundings as either heat or work. For example, when gasoline burns in a car engine, some energy is transformed into work (moving the car), while some is converted to heat (making the engine hot).

The relationship between the change in energy (ΔE), heat (q), and work (w) is given by:

$$\Delta E = q + w$$

4.1.1 Exothermic and Endothermic Reactions

Exothermic Reactions: Reactions that release energy (heat) to the surroundings, making the surroundings warmer. For example, burning wood or dissolving sodium hydroxide ($NaOH$) in water.

Endothermic Reactions: Reactions that absorb energy (heat) from the surroundings, making the surroundings cooler. For example, dissolving potassium nitrate (KNO_3) in water.

The heat transferred during a reaction at constant temperature and pressure is called the **enthalpy of reaction (ΔH)**:

- **Exothermic Reaction**: ΔH is negative ($\Delta H < 0$).
- **Endothermic Reaction**: ΔH is positive ($\Delta H > 0$).

4.2 Energy Changes in Electrochemistry

4.2.1 Electrochemistry

Electrochemistry is the study of the interconversion of electrical energy and chemical energy. It involves using electricity to drive chemical reactions or generating electricity from chemical reactions.

4.2.2 Electrical Conductivity

Electrical Conductivity: The ability of a substance to conduct electricity. Metals conduct electricity because of their free-moving electrons.

- **Metallic Conductivity**: In metals, electrical conductivity is due to the movement of electrons. Metals have a structure where positive ions are surrounded by a "sea" of mobile electrons, allowing them to conduct electricity efficiently.
- **Electrolytic Conductivity**: In electrolytes, the movement of ions (charged particles) allows for the conduction of electricity.

Metals are generally good conductors of electricity due to the presence of free electrons, while non-metals typically do not conduct electricity. However, graphite (a form of carbon) is an exception among non-metals, as it can conduct electricity due to its unique structure.

Key Concepts:

- **Electrochemistry**: The study of how chemical energy and electrical energy are converted.
- **Electrical Conductivity**: The ability to conduct electricity, with metals typically being good conductors due to free-moving electrons.

4.3. Types of Electrochemical Cells

Electrochemical cells can be broadly classified into two main types: **Galvanic (Voltaic) Cells** and **Electrolytic Cells**.

1. **Galvanic (Voltaic) Cells:**
    - **Definition**: These cells convert chemical energy into electrical energy through spontaneous redox reactions.
    - **Structure**: A galvanic cell typically consists of two half-cells connected by a salt bridge or a porous partition. Each half-cell contains an electrode (either metal or non-metal) immersed in an electrolyte solution.
    - **Working Principle**: In a galvanic cell, oxidation occurs at the anode, releasing electrons, while reduction occurs at the cathode, consuming electrons. The flow of electrons from the anode to the cathode through an external circuit generates electric current.
    - **Examples**: The Daniell cell (Zn-Cu cell), where zinc is oxidized at the anode, and copper is reduced at the cathode.
2. **Electrolytic Cells:**
    - **Definition**: These cells convert electrical energy into chemical energy by driving non-spontaneous redox reactions through an external power source.
    - **Structure**: Similar to galvanic cells, electrolytic cells consist of two electrodes immersed in an electrolyte. However, the anode is connected to the positive terminal, and the cathode is connected to the negative terminal of the power source.
    - **Working Principle**: In an electrolytic cell, an external voltage is applied to force electrons to move from the cathode to the anode, driving the redox reaction in a direction that would not occur spontaneously.

- **Examples**: Electrolysis of water, where water is decomposed into hydrogen and oxygen gases by passing an electric current through it.

4.4. Applications of Electrochemical Cells

Electrochemical cells have a wide range of practical applications in various fields. Below are some of the most common uses:

1. **Batteries:**
   - **Primary Cells**: These are non-rechargeable batteries that convert chemical energy into electrical energy until the reactants are exhausted. Examples include dry cells (used in flashlights and remote controls) and alkaline batteries.
   - **Secondary Cells**: These are rechargeable batteries that can be recharged by reversing the chemical reaction. Examples include lead-acid batteries (used in vehicles), lithium-ion batteries (used in smartphones and laptops), and nickel-cadmium batteries.
2. **Fuel Cells:**
   - **Definition**: Fuel cells are electrochemical cells that generate electricity by continuously converting the chemical energy of a fuel (e.g., hydrogen) and an oxidizing agent (e.g., oxygen) into electrical energy.
   - **Applications**: Fuel cells are used in various applications, including powering electric vehicles, portable power sources, and backup power systems for buildings.
3. **Corrosion Prevention:**
   - **Cathodic Protection**: This is a technique used to prevent corrosion of metal structures, such as pipelines and ships, by making the metal the cathode of an electrochemical cell. A sacrificial anode, usually made of a more easily oxidized metal, is connected to the metal structure to provide protection.
4. **Electroplating:**
   - **Definition**: Electroplating is a process in which a metal is deposited onto a surface by passing a current through an electrolytic cell. It is commonly used to coat objects with a thin layer of a different metal, often for decorative or protective purposes.
   - **Applications**: Electroplating is used in various industries, such as jewelry manufacturing, automotive parts production, and electronics, to enhance the appearance and durability of metal surfaces.