CHAPTER III : *Redox Reactions*

2. Electrochemical Cells

An electrochemical cell is a device that produces electrical energy through redox reactions between two electrodes in separate half-cells, connected by a conductive solution, often a salt bridge. In each half-cell, the electrode reacts with ions to produce or consume electrons, which generates an electric current. There are two types :

2.1.Galvanic cell (battery) :

A galvanic cell, also known as a battery, is an electrochemical cell that generates electrical energy through spontaneous redox reactions. In this cell, electrons flow from the anode (where oxidation occurs) to the cathode (where reduction occurs).

2.2. Electrolytic cell :

It is an electrochemical cell where an external electrical energy source drives a non-spontaneous redox reaction. In this cell, electrical energy is supplied to force electrons to move in a direction opposite to that in a galvanic cell. Here, the positive electrode, called the anode, is where oxidation takes place, while the negative electrode, the cathode, is where reduction occurs.

The generator thus formed has two poles:

Positive pole (+) : The cathode, where reduction takes place.

$$Ox_1 + n_1e^- = Red_1$$

Negative pole (-) : The anode, where oxidation occurs :

$$Red_2 = Ox_2 + n2e^-$$

Spontaneous reaction :

$$n_2Ox_1 + n_1Red_2 = n_2Red_1 + n_1Ox_2$$

3. Daniell Cell

This cell was invented by British chemist John Daniell in 1836. It consists of an anode (a zinc strip immersed in a $ZnSO_4$ solution) and a cathode (a copper strip immersed in a $CuSO_4$ solution).

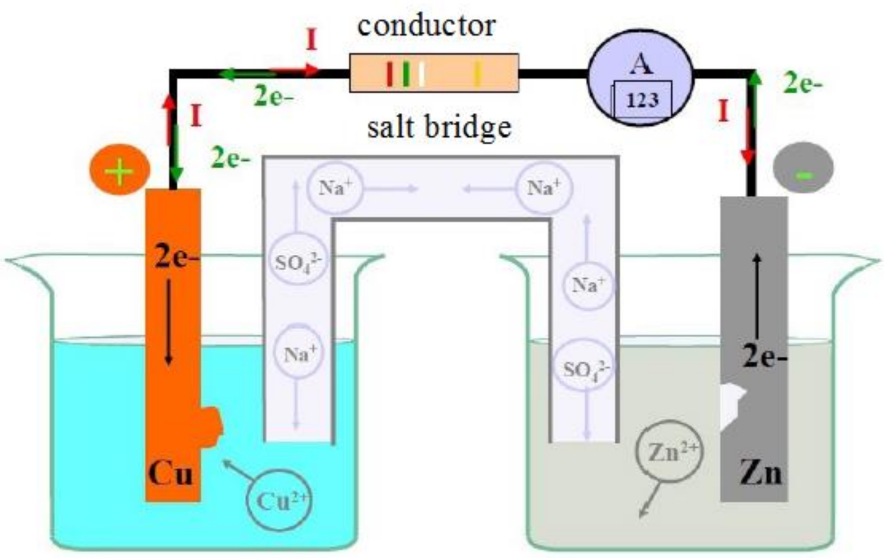

Fig 1. Shematic of a daniell cell

3.1. Operation of the Daniell cell

The two electrodes are linked by an external circuit made up of a resistive conductor in series with an ammeter. The ammeter measures the current that flows from the copper electrode (the positive terminal) to the zinc electrode (the negative terminal). This can also be described as electrons departing from the zinc electrode and traveling toward the copper electrode through the external circuit. Let's now explore the source of this electrical current.

- **Negative Pole (Anode) :** $Zn \rightarrow Zn^{2+} + 2e^-$
- **Positive Pole (Cathode) :** $Cu^{2+} + 2e^- \rightarrow Cu$
- **Overall Reaction :** $Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}$

3.2. Role of the Salt Bridge

- Allows the flow of current within the cell.
- Maintains electrical neutrality in the solutions (by the migration of Na^+ cations in the direction of the electric current within the bridge, and SO_4^{2-} anions in the opposite direction).

The formal representation of this cell is

$$\ominus \; Zn / Zn^{2+} // Cu^{2+} / Cu \; \oplus$$

3.3. Electromotive Force (E.M.F.) of a Battery

The **E.M.F**, denoted as **E**, is the difference between the potentials of the two electrodes, and it is given by:

$$EMF = E = E_{cathode} - E_{anode} = E_+ - E_-$$

Where

- $E_{cathode}$ is the potential of the cathode (positive terminal).
- E_{anode} is the potential of the anode (negative terminal).

3.4. Amount of electricity Q delivered by a battery

When a battery delivers a constant current **I** over a period of time **t**, it circulates an amount of electricity **Q** given by:

$Q = I \times \Delta t$

Where

- Q is the amount of electricity delivered by the battery (**Coulombs**).
- I is the current intensity (**Amperes**).
- Δt is the duration of the battery's operation (**seconds**)

Example

Suppose you have a battery that delivers a current of **2 A** for **4 hours**. We want to calculate the total amount of electricity Q that the battery delivers.

3.5. Battery Capacity

The battery capacity (**Q_{max})**, is the total amount of electric charge that can be transferred from the battery's electrodes during its operation before it reaches its equilibrium state (i.e., when the battery is "used up").

The formula for calculating the maximum amount of electricity is :

$$Q_{\text{max}} = n_{\text{e-max}} \times F$$

Where

- Q_{max} is the maximum amount of electricity (in Coulombs).
- $n_{e\text{-}max}$ is the maximum number of moles of electrons that can be transferred between the electrodes.
- F is Faraday's constant, which is 96,500 C/mol, the charge of one mole of electrons.

Example

Calculate the maximum amount of electricity Qmax delivered by a battery, since the maximum number of moles of electrons that can be transferred $n_{e\text{-}max} = 0.5$ mol