

UNIT 2

(Eco-friendly Portable Energy Convertible Devices)

1. Electrochemistry

Electrochemistry is a branch of chemistry which deals with interconversion of electrical energy to chemical energy and vice versa. For ex: i) In a battery, chemical energy is converted to electrical energy ii) In electroplating / electrolysis electrical energy is converted to chemical energy. Substances that allow electric current to pass through them are known as conductors. For ex: the metals, graphite, fused salts, aqueous solution of acids, bases & salts. While insulator or non-conductor is a substance which does not allow electric current to pass through it. For ex: wood, plastic.

1.1. Electrode Potential: When a metal is kept in a solution of its own ions, the metal acquires either negative or positive charge with respect to the solution. So, a definite potential difference is developed between the metal and the solution. This potential difference is known as electrode potential. For example, when a plate of zinc is placed in a solution having zinc ions (Zn^{2+}), it becomes negatively charged with respect to solution and thus a potential difference is set up between zinc plate and the solution. This potential difference is termed the electrode potential of zinc. Similarly, when copper is placed in a solution having copper ions (Cu^{2+}), it becomes positively charged with respect to solution. A potential difference is set up between the copper plate and the solution. The potential difference thus developed is termed as electrode potential of copper. The potential difference is established due to the formation of electrical double layer known as Helmholtz double layer.

Osmosis is the diffusion solvent molecule through a semipermeable membrane, when a semipermeable membrane separates a solution from a solvent, then only solvent molecules can flow through the semipermeable membrane. During osmosis a pressure is developed on SPM which is known as osmotic pressure.

Solution pressure is a measure of the tendency of molecules or atoms to cross a bounding surface between phases and to enter into a **solution**.

The osmotic pressure of a solution is the pressure difference needed to stop the flow of solvent across a semipermeable membrane. The osmotic pressure of a solution is proportional to the molar concentration of the solute particles in solution. If the solution pressure is greater than the osmotic pressure, the tendency of the metal to lose ions predominates. A potential difference is therefore set up with the metal left with negative charge with respect to the solution. So the positive ions will enter the liquid and the metal negatively charged with respect to solution. The

formation of double layer prevents the further expulsion of ions from the metal, and thus there is rapidly established a state of equilibrium with a definite potential difference, termed as electrode potential.

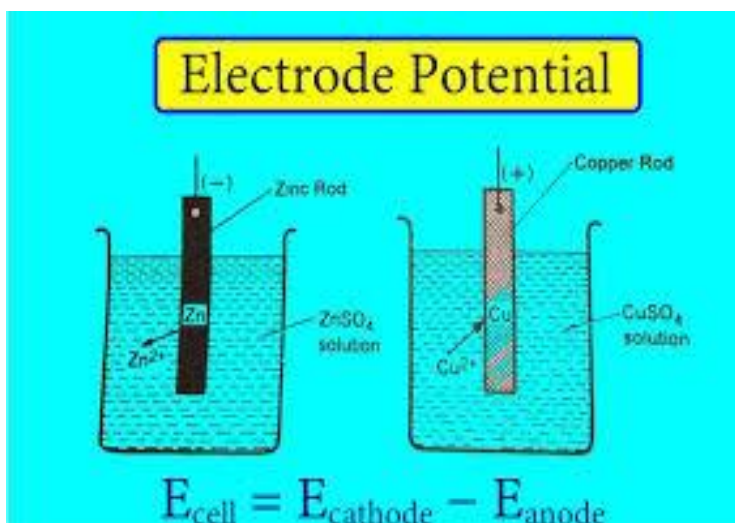


Figure: Representation of positive and negative electrode potential

Following two changes occur when a metal rod is dipped in its salt solution

Oxidation: Metal ions go from the electrode into solution leaving an excess of electrons and thus a negative charge on the electrode.

The conversion of metal atoms into metal ions by the attractive force of polar water molecules.

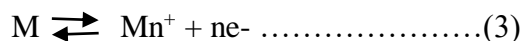


The metal ions go into the solution and the electrons remain on the metal making it negatively charged. The tendency of the metal to change into ions is known as electrolytic solution pressure.

Reduction: Metal ions in solution gain electrons from the electrode leaving a positive charge on the electrode. Metal ions start depositing on the metal outer surface leading to a positive charge on the metal.



In the starting, both these changes occur with different speeds but soon an equilibrium is established.



In practice, one effect is greater than the other,

If first effect is greater than the second, the metal acquires a negative charge with respect to solution and

If the second effect is greater than the first, it acquires positive charge with respect to solution, thus in both the cases a potential difference is build up.

The magnitude of the electrode potential of a metal is a measure of its relative tendency to lose or gain electrons.

Depending on the nature of the metal electrode to lose or gain electrons, the electrode potential may be of two types:

1.1.2.Oxidation potential: When electrode is negatively charged with respect to the solution than it is known as anode. Oxidation occurs at anode.



1.1.3.Reduction potential: When electrode is positively charged with respect to solution than it is known as cathode. Reduction occurs at cathode.



1.2. Standard Electrode Potential

To find out the electrode potentials of different electrodes, it is necessary to specify the concentration of the ions which is present in solution.

The potential difference developed between metal electrode and the solution of its ions at molarity (1M) concentration at 25°C (298 K) is called standard electrode potential. According to the IUPAC convention, the reduction potential alone be called as the electrode potential (EO), i.e., the given value of electrode potential be regarded as reduction potential unless it is specifically mentioned that it is oxidation potential.

Standard reduction potential of an electrode means that reduction reaction is taking place at the electrode. If the reaction is reversed and written as oxidation reaction, the numerical value of electrode potential will remain same, but the sign of standard potential will have to be reversed.

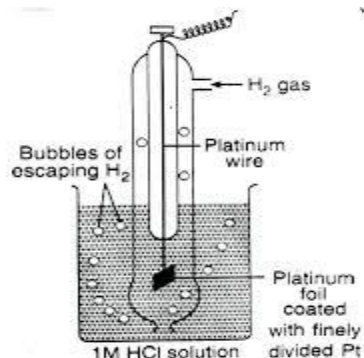


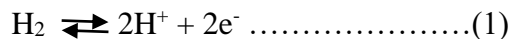
Figure: SHE

1.3. Reference Electrode (Standard Hydrogen Electrode / SHE)

Hydrogen electrode is the primary standard electrode. It consists of a small platinum strip and coated with platinum black (to adsorb hydrogen gas). A platinum wire is surrounded by the platinum strip and sealed in a glass tube as to contact the outer circuit through mercury. The platinum strip and glass tube are surrounded by an outer glass tube which has an inlet for hydrogen gas at the top and a few holes at the base for the escape of excess of hydrogen gas. The

platinum strip is placed in an acid solution having H^+ ion concentration 1 M. Pure hydrogen gas is circulated at 1 atmospheric pressure.

A part of the gas is adsorbed and the rest escapes through holes. This gives an equilibrium between the adsorbed hydrogen and hydrogen ions in the solution.



The temperature of the cell is maintained at $25^\circ C$. By international agreement the standard hydrogen electrode is arbitrarily assigned a potential of exactly ± 0.00 Volt.

The hydrogen electrode thus obtained forms one of two half-cells of a voltaic cell. When this half-cell relates to any other half-cell, a voltaic cell is constituted. The hydrogen electrode can act as cathode or anode with respect to another electrode.

1.4.Electrochemical series: When elements are arranged in increasing order (downwards) of their standard electrode potentials that arrangement is called as electrochemical series.

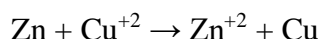
	Half Reaction	Standard Potential (V)
\uparrow stronger oxidizing agent	$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87
	$Pb^{4+} + 2e^- \rightleftharpoons Pb^{2+}$	+1.67
	$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
	$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23
	$Ag^+ + 1e^- \rightleftharpoons Ag$	+0.80
	$Fe^{3+} + 1e^- \rightleftharpoons Fe^{2+}$	+0.77
	$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
	$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
	$Pb^{2+} + 2e^- \rightleftharpoons Pb$	-0.13
	$Fe^{2+} + 2e^- \rightleftharpoons Fe$	-0.44
	$Zn^{2+} + 2e^- \rightleftharpoons Zn$	-0.76
	$Al^{3+} + 3e^- \rightleftharpoons Al$	-1.66
	$Mg^{2+} + 2e^- \rightleftharpoons Mg$	-2.36
	$Li^+ + 1e^- \rightleftharpoons Li$	-3.05
	\downarrow stronger reducing agent	

Features of electrochemical series:

- In electrochemical series a metal with high reduction potential has a higher tendency to undergo reduction.
- A metal having low reduction potential tend to oxidize more easily.

For example standard reduction potential of F_2 / F^- is the highest, so F_2 is easily reduced to F^- . and standard reduction potential of Li^+ / Li is least, so Li^+ is reduced with great difficulty to Li .

- In electrochemical series the metals which are placed on top displace the metals below them from their salt solution. For example: Zn will displace Cu from the solution of Cu^{2+}



- For spontaneous reaction EMF should be positive.
- If EMF is negative value then the reaction is non-spontaneous.

1.5. Cell Potential

The cell potential is the measure of the potential difference between the two half cells in an electrochemical cell. The potential difference is caused by the ability of electrons to flow from one half cell to the other.

- It is not possible to measure the absolute value of the single electrode potential directly. Only the difference in potential between two electrodes can be measured experimentally. So it is necessary to join the electrode with other electrode whose potential is known. This electrode is termed as reference electrode. The EMF of the resulting cell is measured experimentally.
- The EMF of the cell is equal to the sum of potentials on the two electrodes.
- E.M.F. of the cell = $E_{\text{Anode}} + E_{\text{Cathode}}$ = Oxidation potential of anode + Reduction potential of cathode
- It is not possible to measure the absolute value of the single electrode potential directly. Only the difference in potential between two electrodes can be measured experimentally. It is, therefore, necessary to couple the electrode with another electrode whose potential is known. This electrode is termed as reference electrode. The EMF of the resulting cell is measured experimentally.

1.6. Electrochemical cell: The devices which convert chemical energy to electrical energy & electrical energy into chemical energy are known as electrochemical cells they contain two electrodes in contact with an electrolyte, they are mainly of two types.

1.6.1. Galvanic cells: It is an electrochemical cell in which the free energy of chemical reaction is converted into electrical energy i.e. electricity is produced from a spontaneous chemical reaction. Example: voltaic cell or Daniel cell

1.6.2. Electrolytic cell: It is an electrochemical cell in which external electrical energy is used to carry out a non- spontaneous chemical reaction. Examples of electrolysis are the decomposition of water into hydrogen and oxygen, and bauxite into aluminium and other chemicals.

1.7. Galvanic cell : Daniel cell is an example of galvanic cell. It consists of Zn rod and Cu rod; Zn rod is dipped in ZnSO_4 solution and Cu rod dipped in CuSO_4 solution. Each electrode in its electrolytic solution is known as half-cell. The two solutions are connected by U shaped tube known as salt bridge. Thus two electrolytic solutions are in contact with each other. A salt bridge is a U- shaped tube containing concentrated solution of an inert electrolyte like KCl, KNO_3 , K_2SO_4 in order to complete the circuit.

Application of salt bridge:

- Salt bridge prevents the mechanical flow of solution from one-half cell to another.
- It minimizes the liquid-liquid junction potential (Potential arises between two solutions when they are in contact).
- To maintain the electron neutrality.

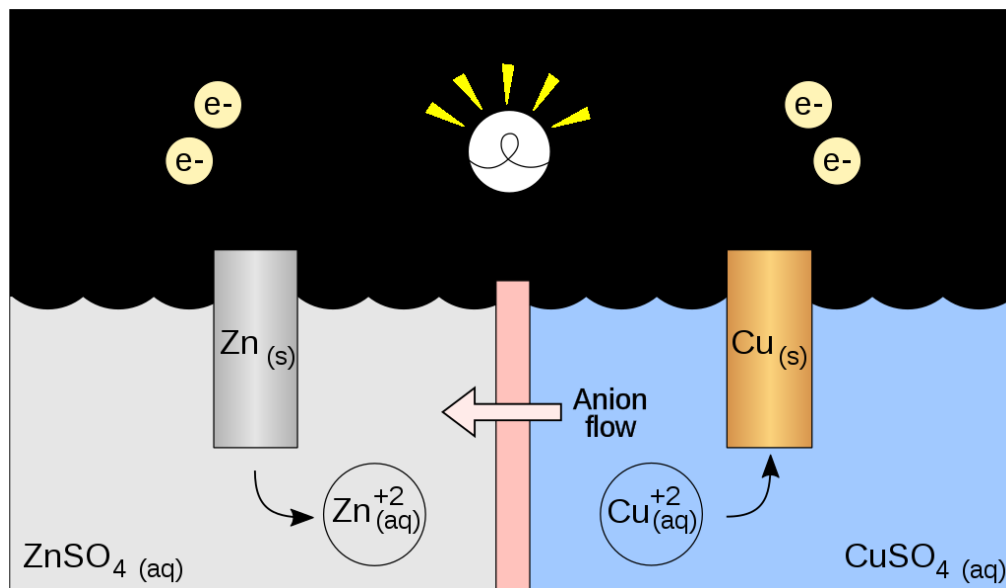
Cell reactions:

The electrode reactions of Daniel Cell are :

At anode: $\text{Zn} \rightarrow \text{Zn}^{+2} + 2\text{e}^-$ (oxidation)(1)

At cathode: $\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu(s)}$ (reduction)(2)

Total cell reaction: $\text{Zn} + \text{Cu}^{+2} \rightarrow \text{Zn}^{+2} + \text{Cu}$ (3)



Green Batteries and their applications

Green batteries refer to environmentally friendly energy storage solutions designed to reduce environmental impact. They typically use sustainable materials and processes, aiming for lower toxicity and better recyclability compared to traditional batteries.

A green battery is first a battery that stores “green” electrons, those generated by renewable sources such as wind or solar. Batteries contain minerals that are mined from Earth's crust, which, like fossil fuels, are not naturally replenished.

Types of Green Batteries

1. **Lithium-Ion Batteries (with Sustainable Practices)**
 - **Description:** While traditional lithium-ion batteries have environmental concerns, improvements in mining practices and recycling technologies are making them greener.
 - **Applications:** Electric vehicles (EVs), renewable energy storage, consumer electronics, and grid stabilization.
2. **Sodium-Ion Batteries**
 - **Description:** Made from abundant sodium, these batteries are being developed as a cost-effective and sustainable alternative to lithium-ion batteries.
 - **Applications:** Similar to lithium-ion batteries, including grid storage and EVs, particularly in regions with abundant sodium resources.
3. **Solid-State Batteries**
 - **Description:** Use a solid electrolyte instead of liquid, potentially improving safety and energy density.
 - **Applications:** Electric vehicles, consumer electronics, and portable devices, offering longer lifespans and quicker charging times.
4. **Flow Batteries**

- **Description:** Store energy in liquid electrolytes housed in external tanks, allowing for scalability and long discharge times.
 - **Applications:** Large-scale renewable energy storage (e.g., solar and wind), grid energy management, and off-grid energy systems.
5. **Organic Batteries**
- **Description:** Utilize organic materials for electrodes, which can be derived from renewable sources.
 - **Applications:** Energy storage for renewable energy systems and potentially consumer electronics, promoting a circular economy.
6. **Biodegradable Batteries**
- **Description:** Made from biodegradable materials that minimize environmental impact when disposed of.
 - **Applications:** Consumer electronics, medical devices, and IoT devices where short-term use is required.

Applications of Green Batteries

1. **Electric Vehicles (EVs)**
 - Green batteries power EVs, reducing reliance on fossil fuels and lowering greenhouse gas emissions. Advances in sodium-ion and solid-state technologies are paving the way for more sustainable options in this sector.
2. **Renewable Energy Storage**
 - Green batteries, particularly flow batteries and lithium-ion batteries, store excess energy generated from renewable sources like solar and wind, helping to balance supply and demand on the grid.
3. **Consumer Electronics**
 - Laptops, smartphones, and other devices benefit from advances in lithium-ion and organic batteries, with a focus on improved recyclability and sustainable sourcing of materials.
4. **Grid Stabilization**
 - Large-scale battery systems (like flow batteries) help maintain grid stability by storing energy during low demand and releasing it during peak demand periods, facilitating the integration of renewable energy.
5. **Off-Grid Systems**
 - Green batteries are ideal for remote or off-grid applications, providing reliable energy storage for homes and businesses powered by renewable sources.
6. **Medical Devices**
 - Biodegradable and organic batteries offer safe and sustainable options for medical devices that require short-term energy solutions.
7. **Wearable Technology**

- Lightweight and efficient green batteries can enhance the performance and sustainability of wearable tech, such as fitness trackers and smartwatches.