GALVANIC CELL AND ELECTRODE POTENTIAL

Galvanic Cell or Voltaic Cell

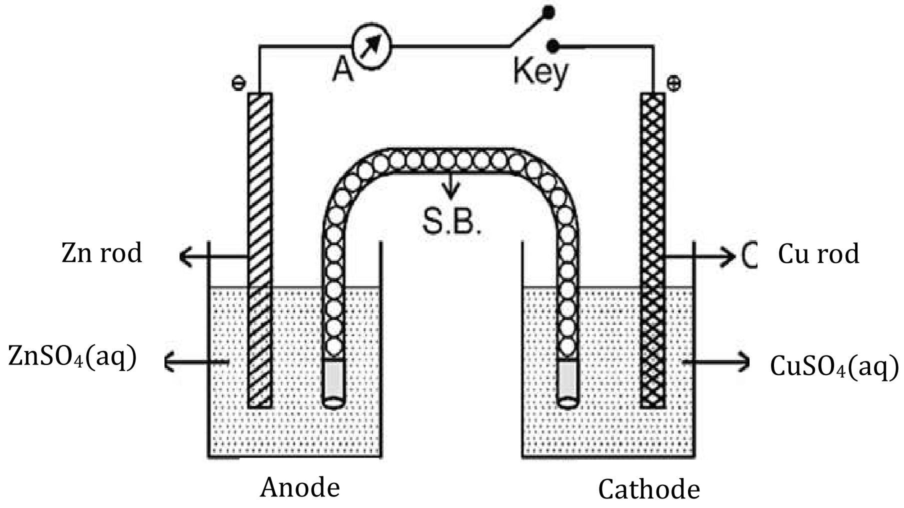
Redox Couple

A galvanic cell, also known as a voltaic cell, is a straightforward electrochemical cell that utilizes a redox reaction to convert chemical energy into electrical energy. In essence, it allows the generation of electricity through a redox reaction wherein oxidation and reduction occur in separate compartments. Each compartment is equipped with a metallic conductor, typically in the form of a metal rod, immersed in a suitable electrolytic solution of the same metal, and the metallic rod serves as an electrode.

The compartment with the electrode immersed in the electrolyte solution is referred to as a half-cell, and each half-cell contains a redox couple.

A redox couple is a solution containing both the reduced and oxidized forms of a substance, participating in either the oxidation or reduction half-reaction. This is symbolized as M^{+n}/M, where M represents the oxidized form, and M^{+n} denotes the reduced form.

To assemble a galvanic cell, two half-cells are externally connected by a conducting wire and internally connected through a salt bridge. The salt bridge is a U-tube filled with an electrolyte solution, such as KCl or KNO_3, solidified by boiling and subsequently cooling with Agar-Agar. Agar-Agar maintains the salt bridge in a semi-solid, jelly-like form, facilitating the connection between the two half-cells and providing electrical contact without mixing their solutions. Additionally, the salt bridge ensures the electrical neutrality of the solutions in the two half-cells.



Daniel or Voltaic Call

Externally, a metallic wire is connected to an ammeter and a switch, creating a connection between the two electrodes. In this setup, conventionally, the left-hand side electrode serves as the anode, where oxidation occurs, while the right-hand side electrode functions as the cathode, where reduction takes place. The anode, being electron-rich, is assigned a negative sign, whereas the cathode, being electron-deficient, is given a positive sign. Electrons flow externally from the anode to the cathode, and the direction of the current flow is considered from the cathode to the anode internally.

Anodic oxidation $Zn_{(s)} \rightarrow Zn^{+2}_{(aq)} + 2e^-$

Cathodic reduction $Cu^{+2}_{(aq)} + 2e^- \rightarrow Cu_{(S)}$

Net Cell reaction $Zn_{(s)} + Cu^{+2}_{(aq)} \rightarrow Nn^{+2}_{(aq)} + Cu_{(s)}$

The representation of this cell can be succinctly expressed in a single line, referred to as cell notation.

i.e., $Zn_{(s)} \mid Zn^{+2}_{(aq)}, C_1 \,||\, Cu^{+2}_{(aq)}, C_2 \mid Cu_{(s)}$

The internal flow of electricity between solutions in two separate beakers occurs through the migration of ions facilitated by a salt bridge. The external flow of current is viable only in the presence of a potential difference between the two electrodes.

To comprehend potential difference, one must grasp the concept of electrode potential. This is defined as a measure of the relative tendency of the active species in a process to either remain in an oxidized or reduced form. It can also be expressed as the capability or tendency of an electrode to either lose or gain electrons.

When the concentration of each species involved in the electrode reaction is unity (or, for any gas involved, its pressure is one atmosphere), and the reaction takes place at 298 K, the potential of each electrode is termed the standard electrode potential ($E°$) Conventionally, $E°_{red}$ for the hydrogen electrode is considered 0.00 V.

A comparative list of $E°_{red}$ values for various redox couples has been compiled with reference to the standard hydrogen electrode, forming what is known as the electrochemical series. This series holds great significance in electrochemistry.

In the context of $E°$ values:

- A negative value indicates that the redox couple is a stronger reducing agent than the $\frac{H^+}{H_2}$ couple.
- A positive value indicates that the redox couple is a weaker reducing agent than the $\frac{H^+}{H_2}$ couple.

The cell potential ($E°_{cell}$) is calculated as the difference between the standard reduction potentials of the cathode ($E°_{cathode}$) and anode ($E°_{anode}$). It is important to note that both standard potentials must be reduction potentials.

Furthermore, for any electrode, the standard reduction potential ($E°_{red}$) is equivalent to the negative of the standard oxidation potential ($E°_{oxid}$).