

Chapter- 4

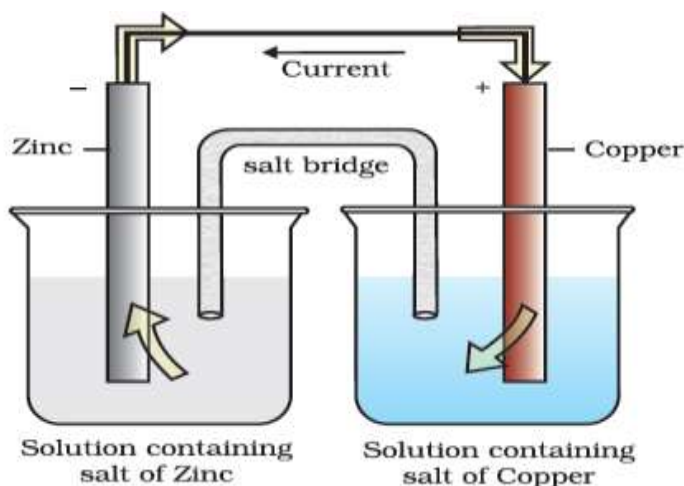
ELECTROCHEMISTRY

The branch of physical chemistry which deals with the study of conversion of electrical energy into chemical energy and vice versa is called electrochemistry.

Conversion of electrical energy into chemical energy occurs in electrolytic cell and the conversion of chemical energy into electrical energy occurs in electrochemical cell or voltaic cell.

Electrochemical Cell / Galvanic cell:

- a) A galvanic cell is an electrochemical cell that converts the chemical energy of a spontaneous redox reaction into electrical energy. In this device the Gibbs energy of the spontaneous redox reaction is converted into electrical work which may be used for running motor or other electrical gadgets like heater, fan, geyser etc. Ex: Daniel cell or Galvanic cell.
- b) The electrochemical cell consists of two half cells such as; oxidation half cell and reduction half cell. When oxidation and reduction takes place in two separate beakers, then chemical energy is converted into electrical energy. This cell has an electrical potential equal to 1.1 V when concentration of Zn^{2+} and Cu^{2+} ions is unity (1 mol dm^{-3}). Such a device is also called **voltaic cell**.



c) The two half-cells are connected by a metallic wire through a voltmeter and a switch externally and the electrolytes of the two half-cells are connected internally through a salt bridge. Electrons flow through the wire and ions flow through the salt bridge to induce current in the cell.

d) **In Zn-CuSO₄ cell.**

In left side, Zn is oxidised to Zn²⁺ ion and in right side Cu²⁺ from solution is reduced and deposited on copper electrode by receiving electron through wire. As a result concentration of Zn²⁺ ion in left side beaker and concentration of SO₄²⁻ ion in right side beaker is increased. The excess of ions flow through the salt bridge to maintain electrical neutrality and to generate current in the cell.

e) **Function of Salt bridge :**

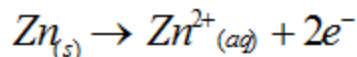
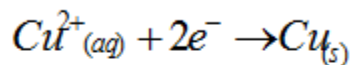
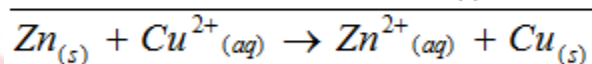
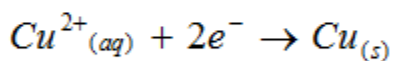
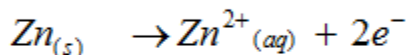
Salt bridge is a U-shaped tube containing inert electrolyte. The electrolyte is a semisolid paste of either KCl, KNO₃, K₂SO₄ or NH₄Cl in gelatine or agar-agar gelatine solution etc.

- (1) Salt bridge helps to maintain electrical neutrality between the two half cells.
- (2) It doesn't react with the electrolytes present in two half cells.
- (3) It allows only flow of ions through it. I.e. it complete the cell circuit
- (4) Salt bridge completes the circuit to generate emf in the cell.

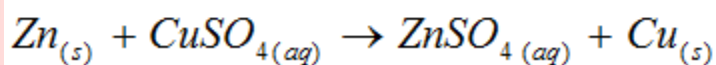
In electrochemical cell, electrons flow from high oxidation potential to low oxidation potential and the current flows from low oxidation potential to high oxidation potential.

- Electrons flow from H.O.P to L.O.P or L.R.P to H.R.P
- Current flows from L.O.P. to H.O.P or H.R.P to L.R.P

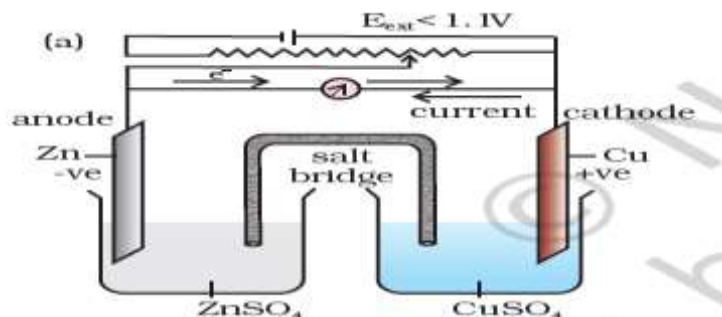
In electrochemical cell where oxidation takes place is represented as anode and where reduction takes place is represented as cathode. Anode is assigned as -ve sign and cathode as +ve sign.

g) Cell reaction:1) At anode (Oxidation half reaction):2) At cathode (Reduction half reaction)3) Full cell reaction (Oxidation half + Reduction half)

It is also written as:

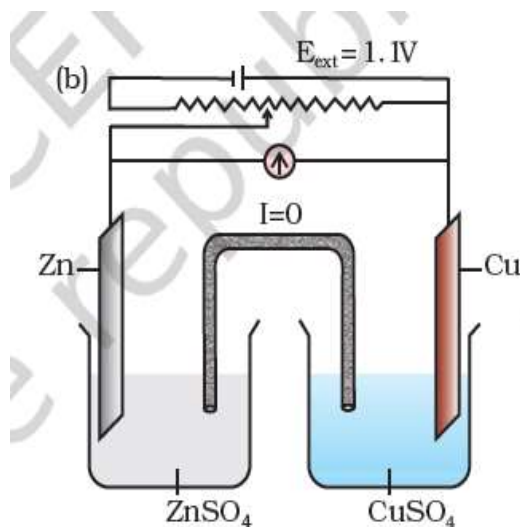
h) Applying external potential:

- 1) If an external opposite potential is applied in the galvanic cell and increased slowly, then the reaction continues to take place till the opposing voltage reaches the value 1.1V when, the reaction stops altogether and no current flows through the cell.
- 2) Any further increase in the external potential again starts the reaction but in the opposite direction. It now functions as an **electrolytic cell**, a device for using electrical energy to carry non-spontaneous chemical reactions



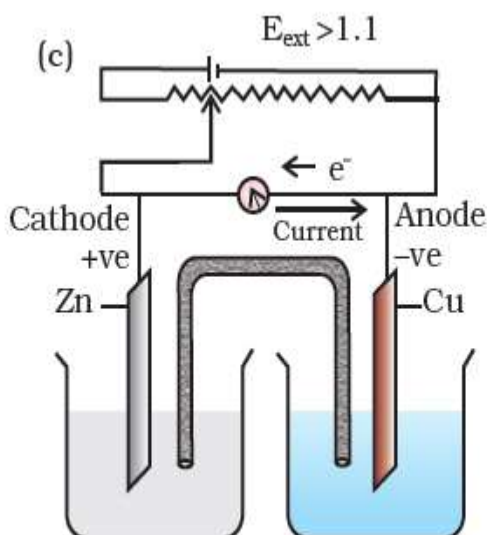
When $E_{ext} < 1.1 \text{ V}$

- (i) Electrons flow from Zn rod to Cu rod hence current flows from Cu to Zn.
- (ii) Zn dissolves at anode and copper deposits at cathode.



When $E_{\text{ext}} = 1.1 \text{ V}$

- (i) No flow of electrons or current.
- (ii) No chemical reaction.



When $E_{\text{ext}} > 1.1 \text{ V}$

- (i) Electrons flow from Cu to Zn and current flows from Zn to Cu.
- (ii) Zinc is deposited at the zinc electrode and copper dissolves at copper electrode.

Electrode Potential:-

It may be defined as the tendency of a metal to lose or gain electrons when it is placed in contact with its own ions. When the concentrations of all the species involved in a half-cell is unity then the electrode potential is known as **standard electrode potential**. According to IUPAC convention, standard reduction potentials are now called standard electrode potentials. The electrode potential will be named as oxidation or reduction potential depending upon whether oxidation or reduction takes place.

Oxidation potential is the tendency of an electrode to lose electrons. The electrode having higher oxidation potential can lose electron easily or oxidised easily and possesses higher reducing power.

Reduction potential is the tendency of an electrode to gain electrons. The electrode having higher reduction potential is easily reduced and possesses higher oxidising power.

As oxidation and reduction are opposite to each other thus

$$\text{Reduction potential} = - \text{Oxidation potential}$$

Or

$$\text{Oxidation potential} = - \text{Reduction potential.}$$

$$\text{Cu}^{2+} / \text{Cu} = +0.34V \text{ (Reduction potential)}$$

$$\text{Cu} / \text{Cu}^{2+} = -0.34V \text{ (Oxidation potential)}$$

Cell voltage / Cell Potential / E.M.F of cell (E° cell)

The potential difference between the two electrodes of a galvanic cell is called the *cell potential* and is measured in volts. The *cell potential* is the difference between the electrode potentials (reduction potentials) of the cathode and anode.

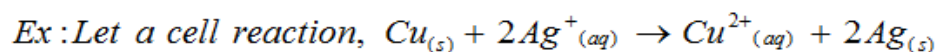
E.M.F of cell (E°_{cell}) = Oxidation potential at left electrode + Reduction potential at right electrode

$$= \text{Oxidation potential at anode} + \text{Reduction potential at cathode}$$

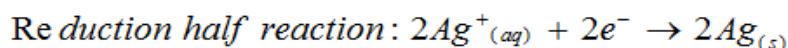
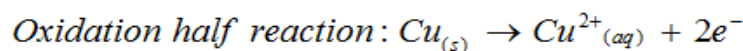
$$= - \text{Reduction potential at anode} + \text{Reduction potential at cathode.}$$

$$= \text{Reduction potential at cathode} - \text{Reduction potential at anode}$$

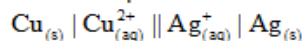
$$\Rightarrow E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

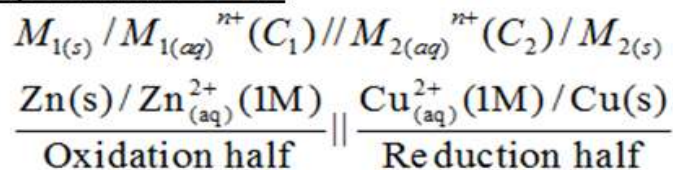


Half cell reactions :

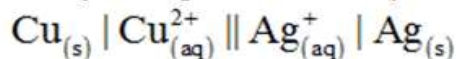


The cell can be represented as

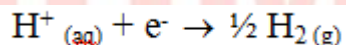


LECT-2**f) Representation of Cell:**

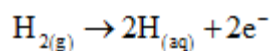
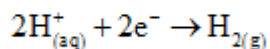
Similarly Cu-Ag cell can be represented as

**Measurement of Electrode Potential (Standard hydrogen electrode):**

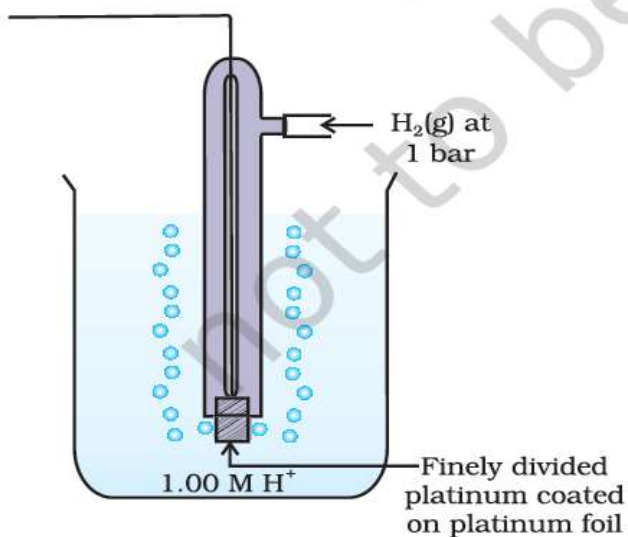
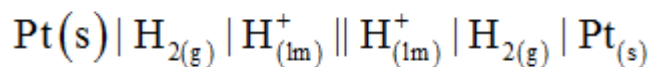
The potential of individual half-cell cannot be measured. We can measure only the difference between the two half-cell potentials that gives the emf of the cell. If we arbitrarily choose the potential of one electrode (half-cell) then that of the other can be determined with respect to this. According to convention, a half-cell **called standard hydrogen electrode** represented by $\text{Pt}_{(s)} / \text{H}_2 (g) / \text{H}^{+}_{(aq)}$, is assigned a zero potential at all temperatures corresponding to the reaction.



Standard hydrogen electrode is a reference electrode and its reduction potential is arbitrarily assigned as zero volt at all temperature. A S.H.E consists of a platinum electrode coated with platinum black. The electrode is dipped in acidic solution having 1M concentration of H^{+} ions. Pure hydrogen gas at 1 bar pressure is bubbled through the solution at a temperature 298K. The hydrogen electrode can act both as anode as well as cathode.

When SHE act as anode:-**When SHE act as cathode:-**

Symbolic representation of SHE.



Standard electrode potentials (E^0) of the electrodes are calculated By using SHE .If the concentrations of the oxidised and the reduced forms of the species in the right hand half-cell are unity, then the cell potential is equal to standard electrode potential, E^0 of the given half-cell.

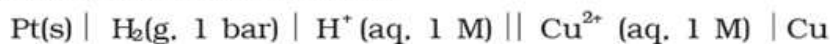
Changing your Tomorrow ▲

$$E^{\ominus} = E^{\ominus}_R - E^{\ominus}_L$$

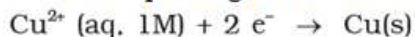
As E^{\ominus}_L for standard hydrogen electrode is zero.

$$E^{\ominus} = E^{\ominus}_R - 0 = E^{\ominus}_R$$

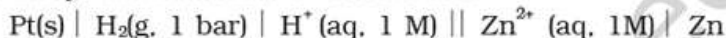
The measured emf of the cell:



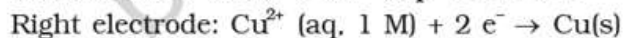
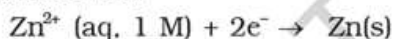
is 0.34 V and it is also the value for the standard electrode potential of the half-cell corresponding to the reaction:



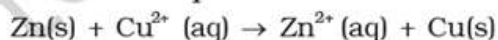
Similarly, the measured emf of the cell:



is -0.76 V corresponding to the standard electrode potential of the half-cell reaction:



The overall reaction of the cell is the sum of above two reactions and we obtain the equation:



If the standard electrode potential of an electrode is greater than zero then its reduced form is more stable compared to hydrogen gas. Similarly, if the standard electrode potential is negative then hydrogen gas is more stable than the reduced form of the species.

$$\begin{aligned} \text{emf of the cell} &= E^{\ominus}_{\text{cell}} = E^{\ominus}_R - E^{\ominus}_L \\ &= 0.34\text{V} - (-0.76)\text{V} = 1.10 \text{ V} \end{aligned}$$

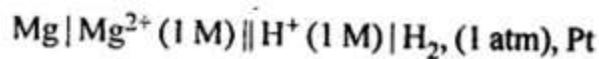
Intext Questions:

- How would you determine the standard electrode potential of the system $\text{Mg}^{2+} \mid \text{Mg}$?
- Can you store copper sulphate solutions in a zinc pot?
- Consult the table of standard electrode potentials and suggest three substances that can oxidise ferrous ions under suitable conditions.

Solution:

3.1. How would you determine the standard electrode potential of the system $\text{Mg}^{2+} \mid \text{Mg}$?

Ans: A cell will be set up consisting of $\text{Mg}/\text{MgSO}_4(1 \text{ M})$ as one electrode and standard hydrogen electrode $\text{Pt}, \text{H}_2(1 \text{ atm})/\text{H}^+(1 \text{ M})$ as second electrode, measure the EMF of the cell and also note the direction of deflection in the voltmeter. The direction of deflection shows that e^- 's flow from mg electrode to hydrogen electrode, i.e., oxidation takes place on magnesium electrode and reduction on hydrogen electrode. Hence, the cell may be represented as follows :



$$E_{\text{cell}}^{\circ} = E_{\text{H}^+ / \frac{1}{2}\text{H}_2}^{\circ} - E_{\text{Mg}^{2+} / \text{Mg}}^{\circ}$$

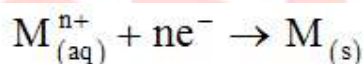
$$\text{Put } E_{\text{H}^+ / \frac{1}{2}\text{H}_2}^{\circ} = 0$$

$$\therefore E_{\text{Mg}^{2+} / \text{Mg}}^{\circ} = -E_{\text{cell}}^{\circ}$$

LECT-3

Nernst Equation:

We have assumed in the previous section that the concentration of all the species involved in the electrode reaction is unity. This need not be always true. Nernst showed that for the electrode reaction:



The electrode potential at any concentration measured with respect to standard hydrogen electrode can be represented by:

$$E_{(\text{M}^{n+} / \text{M})} = E_{(\text{M}^{n+} / \text{M})}^{\ominus} - \frac{RT}{nF} \ln \frac{[\text{M}]}{[\text{M}^{n+}]}$$

but concentration of solid M is taken as unity and we have

$$E_{(\text{M}^{n+} / \text{M})} = E_{(\text{M}^{n+} / \text{M})}^{\ominus} - \frac{RT}{nF} \ln \frac{1}{[\text{M}^{n+}]}$$

Substituting the value of $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$ and $F = 96487 \text{ C mol}^{-1}$ we get.

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{n} \log \frac{1}{[\text{M}^{n+}]} \text{ at } 298\text{k}$$

In Daniel cell, the electrode potential for any given concentration of Cu^{2+} and Zn^{2+} ions, we write

For Cathode:

$$E_{\text{Cu}^{2+} / \text{Cu}} = E_{\text{Cu}^{2+} / \text{Cu}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Cu}^{2+}]}$$

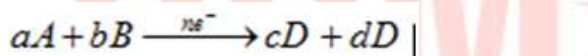
For Anode:

$$E_{\text{Zn}^{2+} / \text{Zn}} = E_{\text{Zn}^{2+} / \text{Zn}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]}$$

The cell potential is calculated as:

$$\begin{aligned}
 E_{\text{cell}} &= E_{\text{Cu}^{2+}/\text{Cu}} - E_{\text{Zn}^{2+}/\text{Zn}} \\
 &= \left[E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Cu}^{2+}]} \right] - \left[E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]} \right] \\
 &= E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Cu}^{2+}]} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} + \frac{0.0591}{2} \log \frac{1}{[\text{Zn}^{2+}]} \\
 &= E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} + \frac{0.0591}{2} \left[\log \frac{1}{[\text{Zn}^{2+}]} - \log \frac{1}{[\text{Cu}^{2+}]} \right] \\
 &= E_{\text{cell}}^{\circ} + \frac{0.0591}{2} \left[\log \frac{[\text{Cu}^{2+}]}{[\text{Zn}^{2+}]} \right] \\
 \Rightarrow E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}
 \end{aligned}$$

In general, if the cell reaction is



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303RT}{nF} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

At 25°C,

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log Q_c$$



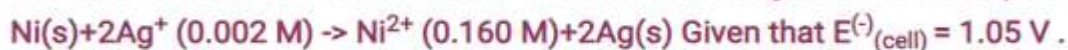
3.4. Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

Ans. For hydrogen electrode, $\text{H}^+ + \text{e}^- \rightarrow 1/2 \text{H}_2$

Applying Nernst equation,

$$\begin{aligned}
 E_{\text{H}^+ / \frac{1}{2}\text{H}_2} &= E_{\text{H}^+ / \frac{1}{2}\text{H}_2}^{\circ} - \frac{0.0591}{n} \log \frac{1}{[\text{H}^+]} \\
 &= 0 - \frac{0.0591}{1} \log \frac{1}{10^{-10}} \\
 &\quad \left\{ \begin{array}{l} \text{pH} = 10 \\ \Rightarrow [\text{H}^+] = 10^{-10} \text{ M} \end{array} \right\} \\
 &= -0.0591 \times 10 \\
 &= -0.591 \text{ V}
 \end{aligned}$$

3.5. Calculate the emf of the cell in which the following reaction takes place:



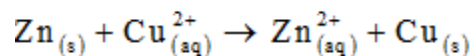
Ans:

Applying Nernst equation,

$$\begin{aligned} E_{\text{cell}} &= E^{\ominus}_{\text{cell}} - \frac{0.0591}{n} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \\ &= 1.05 \text{ V} - \frac{0.0591}{2} \log \frac{0.160}{(0.002)^2} \\ &= 1.05 - \frac{0.0591}{2} \log(4 \times 10^4) \\ &= 1.05 - \frac{0.0591}{2} (4.6021) \\ &= 1.05 - 0.14 \text{ V} \\ &= 0.91 \text{ V} \end{aligned}$$

Equilibrium constant from Nernst Equation:-

Consider the reaction taking place in a Daniel cell.



As time passes, the concentration of Zn^{2+} keeps on increasing while the concentration of Cu^{2+} keeps on decreasing. At the same time voltage of the cell as read on the voltmeter keeps on decreasing. After some time, we shall note that there is no change in the concentration of Cu^{2+} and Zn^{2+} ions and at the same time, voltmeter gives zero reading. This indicates that equilibrium has been attained. In this situation the Nernst equation may be written as:

$$\begin{aligned} E_{\text{cell}} = 0 &= E^{\ominus}_{\text{cell}} - \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \\ E^{\ominus}_{\text{cell}} &= \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \end{aligned}$$

$$\text{at equilibrium } \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = K_c$$

At $T = 298\text{K}$ the above equation can be written as

$$E_{\text{cell}}^{\ominus} = \frac{0.0591}{2} \log K_c = 1.1\text{V}$$

$$\log K_c = \frac{(1.1\text{V} \times 2)}{0.0591\text{V}} = 37.288$$

$$K_c = 2 \times 10^{37} \text{ at } 298\text{K}$$

$$\text{In general, } E_{\text{cell}}^{\ominus} = \frac{2.303RT}{nF} \log K_c$$

Calculate the equilibrium constant of the reaction:



$$E_{(\text{cell})}^{\ominus} = 0.46\text{ V}$$

$$\text{Solution } E_{(\text{cell})}^{\ominus} = \frac{0.059\text{ V}}{2} \log K_c = 0.46\text{ V or}$$

$$\log K_c = \frac{0.46\text{ V} \times 2}{0.059\text{ V}} = 15.6$$

$$K_c = 3.92 \times 10^{15}$$

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LECT-4

Electrochemical Cell and Gibbs Energy of the Reaction:

Electrical work done in one second is equal to electrical potential multiplied by total charge passed.

$$\Delta_r G = - nFE_{(\text{cell})}$$

If the concentration of all the reacting species is unity, then

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus}$$

So we have

$$\Delta_r G^{\ominus} = -nFE_{\text{cell}}^{\ominus}$$

Standard Gibb's energy of reaction ($\Delta_r G^{\ominus}$) is related to equilibrium constant (K) by the reaction

$$\Delta_r G^{\ominus} = -RT \ln K$$

$$\Rightarrow \Delta_r G^{\ominus} = -2.303RT \log K$$

Q. The standard electrode potential for Daniell cell is 1.1V. Calculate the standard Gibbs energy for the reaction:



Solution $\Delta_r G^\ominus = -nFE_{(\text{cell})}^\ominus$

n in the above equation is 2, $F = 96487 \text{ C mol}^{-1}$ and $E_{(\text{cell})}^\ominus = 1.1 \text{ V}$

$$\begin{aligned} \text{Therefore, } \Delta_r G^\ominus &= -2 \times 1.1 \text{ V} \times 96487 \text{ C mol}^{-1} \\ &= -21227 \text{ J mol}^{-1} \\ &= -212.27 \text{ kJ mol}^{-1} \end{aligned}$$

Conductance of Electrolytic Solutions:

(c) Specific resistance/Resistivity (ρ) :- The resistance of a conductor is directly proportional to the length (ℓ) and area of cross section (A) of conductor.

So,

$$R \propto \ell, R \propto \frac{\ell}{A}$$

$$\Rightarrow R \propto \frac{\ell}{A} \Rightarrow R = \rho \frac{\ell}{A}$$

ρ is called specific resistance or resistivity of conductor.

$$\Rightarrow \rho = \frac{RA}{\ell}$$

When $\ell = 1\text{cm}$, $A = 1\text{cm}^2$ $\rho = R$

Hence, specific resistance is defined as the resistance offered by a conductor of 1cm^3 volume.

Unit of ρ :- Ohm cm or Ohm m.

(d) Specific Conductance (Conductivity)

It is the reciprocal of specific resistance

(b) Conductance (G) :- It is the reciprocal of resistance.

$$G = \frac{1}{R} = \frac{1}{\rho \frac{l}{A}} = k \frac{A}{l}$$

Where k = Conductivity

$$K = \frac{1}{\rho} = \frac{1}{\frac{RA}{l}}$$

$$\Rightarrow K = \frac{1}{R} \times \frac{l}{A} \Rightarrow K = G \times G^* \text{ where } \frac{l}{A} = \text{cell constant } (G^*).$$

Unit of conductivity (K) is $\text{ohm}^{-1}\text{cm}^{-1}$ or $\text{ohm}^{-1}\text{m}^{-1}$

Electrical conductance through metals is called metallic or electronic conductance and is due to the movement of electrons. The electronic conductivity depends on.

- (a) The nature and structure of the metal
- (b) The number of valence electrons per atom
- (c) Temperature (it decreases with the increases of temperature)

The conductance of electricity by ions present in the solution is called electrolytic or ionic conductance. The conductance of ionic solution depends on.

- (a) The nature of the electrolyte added.
- (b) Size of the ion produced and their salvation
- (c) Nature of the solvent and their viscosity
- (d) Concentration of the electrolyte
- (e) Temperature it increases with the increase of temperature.

3.9 The resistance of a conductivity cell containing 0.001 M KCl solution at 298 K is 1500 Ω . What is the cell constant if conductivity of 0.001 M KCl solution at 298 K is $0.146 \times 10^{-3} \text{ S cm}^{-1}$?

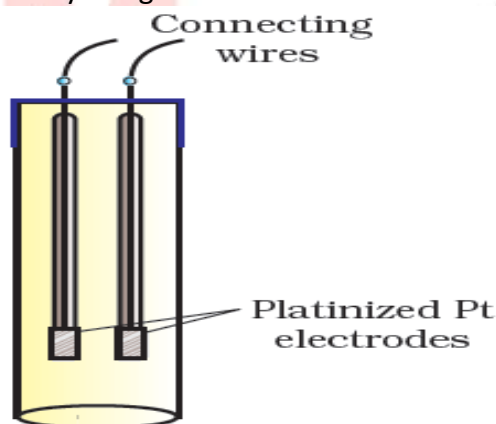
Sol.

$$\begin{aligned} \text{Cell constant} &= \frac{\text{Conductivity}}{\text{Conductance}} \\ &= \text{Conductivity} \times \text{Resistance} \\ &= 0.146 \times 10^{-3} \text{ S cm}^{-1} \times 1500 \Omega \\ &= 0.219 \text{ cm}^{-1}. \end{aligned}$$

LECT-5

Measurement of the Conductivity of Ionic Solutions:

We know that accurate measurement of an unknown resistance can be performed on a Wheatstone bridge. However, for measuring the resistance of an ionic solution we face two problems. Firstly, passing direct current (DC) changes the composition of the solution. Secondly, a solution cannot be connected to the bridge like a metallic wire or other solid conductor. The first difficulty is resolved by using an alternating current (AC) source of power. The second problem is solved by using a specially designed vessel called **conductivity cell**.



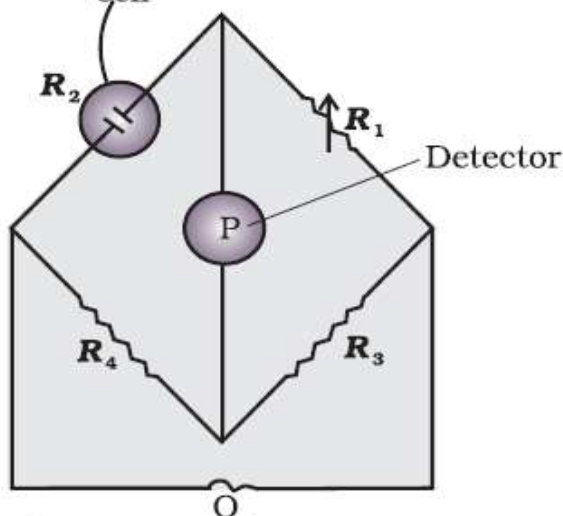
Basically it consists of two platinum electrodes coated with platinum black (finely divided metallic Pt is deposited on the electrodes electrochemically). These have area of cross section equal to 'A' and are separated by distance 'l'. Therefore, solution confined between these electrodes is a column of length l and area of cross section A. The resistance of such a column of solution is then given by the equation:

$$R = \rho \frac{l}{A} = \frac{l}{\kappa A}$$

whose conductivity is known accurately at various concentrations. And at different temperatures The quantity l/A is called cell constant denoted by the symbol, G^* . It depends on the distance between the electrodes and their area of cross-section and has the dimension of length⁻¹ and can be calculated if we know l and A. Measurement of l and A is not only inconvenient but also unreliable. The cell constant is usually determined by measuring the resistance of the cell containing a solution whose conductivity is already known. For this purpose, we generally use KCl solutions. The cell constant, G^* , is then given by the equation:

$$G^* = \frac{l}{A} = R \kappa$$

Conductivity
cell



(Wheatstone bridge)

Once the cell constant is determined, we can use it for measuring the resistance or conductivity of any solution. It consists of two resistances R_3 and R_4 , a variable resistance R_1 and the conductivity cell having the unknown resistance R_2 . The Wheatstone bridge is fed by an oscillator O (a source of a.c. power in the audio frequency range 550 to 5000 cycles per second). P is a suitable detector (a headphone or other electronic device) and the bridge is balanced when no current passes through the detector. Under these conditions:

$$\text{Unknown resistance } R_2 = \frac{R_1 R_4}{R_3}$$

Once the cell constant and the resistance of the solution in the cell is determined, the conductivity of the solution is given by the equation:

$$\kappa = \frac{\text{cell constant}}{R} = \frac{G^*}{R}$$

(e) Molar conductivity' (\wedge_m):- It is defined as the conducting power of all the ions produced by one mole of electrolytic solution. Kept between the electrodes of a conductivity cell at unit distance but having area of cross section large enough to accommodate sufficient volume of solution that contain one mole of electrolyte.

$$\text{Molar conductivity } \wedge_m = \frac{k}{C}$$

In the above equation, if κ is expressed in S m^{-1} and the concentration, c in mol m^{-3} then the units of Λ_m are in $\text{S m}^2 \text{mol}^{-1}$. It may be noted that:
 $1 \text{ mol m}^{-3} = 1000(\text{L/m}^3) \text{ molarity (mol/L)}$, and hence

$$\Lambda_m (\text{S cm}^2 \text{mol}^{-1}) = \frac{\kappa (\text{S cm}^{-1})}{1000 \text{ L m}^{-3} \times \text{molarity (mol L}^{-1}\text{)}}$$

If we use S cm^{-1} as the units for κ and mol cm^{-3} , the units of concentration, then the units for Λ_m are $\text{S cm}^2 \text{mol}^{-1}$. It can be calculated by using the equation:

$$\Lambda_m (\text{S cm}^2 \text{mol}^{-1}) = \frac{\kappa (\text{S cm}^{-1}) \times 1000 (\text{cm}^3 / \text{L})}{\text{molarity (mol / L)}}$$

Q. Resistance of a conductivity cell filled with 0.1 mol L^{-1} KCl solution is 100 ohm. If the resistance of the same cell when filled with 0.02 mol L^{-1} KCl solution is 520 ohm, calculate the conductivity and molar conductivity of 0.02 mol L^{-1} KCl solution. The conductivity of 0.1 mol L^{-1} KCl solutions is 1.29 S/m .

Solution: The cell constant is given by the equation:

Cell constant = $G^* = \text{conductivity} \times \text{resistance}$

= $1.29 \text{ S/m} \times 100 \text{ W} = 129 \text{ m}^{-1} = 1.29 \text{ cm}^{-1}$

Conductivity of 0.02 mol L^{-1} KCl solution = cell constant / resistance

$$= \frac{G^*}{R} = \frac{129 \text{ m}^{-1}}{520 \Omega} = 0.248 \text{ S m}^{-1}$$

$$\begin{aligned} \text{Concentration} &= 0.02 \text{ mol L}^{-1} \\ &= 1000 \times 0.02 \text{ mol m}^{-3} = 20 \text{ mol m}^{-3} \end{aligned}$$

$$\begin{aligned} \text{Molar conductivity} &= \Lambda_m = \frac{\kappa}{c} \\ &= \frac{248 \times 10^{-3} \text{ S m}^{-1}}{20 \text{ mol m}^{-3}} = 124 \times 10^{-4} \text{ S m}^2 \text{mol}^{-1} \end{aligned}$$

$$\text{Alternatively, } \kappa = \frac{1.29 \text{ cm}^{-1}}{520 \Omega} = 0.248 \times 10^{-2} \text{ S cm}^{-1}$$

$$\begin{aligned} \text{and } \Lambda_m &= \kappa \times 1000 \text{ cm}^3 \text{ L}^{-1} \text{ molarity}^{-1} \\ &= \frac{0.248 \times 10^{-2} \text{ S cm}^{-1} \times 1000 \text{ cm}^3 \text{ L}^{-1}}{0.02 \text{ mol L}^{-1}} \\ &= 124 \text{ S cm}^2 \text{mol}^{-1} \end{aligned}$$

Q. The electrical resistance of a column of 0.05 mol L^{-1} NaOH solutions of diameter 1 cm and length 50 cm is 5.55×10^3 ohm. Calculate its resistivity, conductivity and molar conductivity.

Solution:

$$A = \pi r^2 = 3.14 \times 0.5^2 \text{ cm}^2 = 0.785 \text{ cm}^2 = 0.785 \times 10^{-4} \text{ m}^2$$

$$l = 50 \text{ cm} = 0.5 \text{ m}$$

$$R = \frac{\rho l}{A} \quad \text{or} \quad \rho = \frac{RA}{l} = \frac{5.55 \times 10^3 \Omega \times 0.785 \text{ cm}^2}{50 \text{ cm}} = 87.135 \Omega \text{ cm}$$

$$\begin{aligned} \text{Conductivity} = \kappa &= \frac{1}{\rho} = \left(\frac{1}{87.135} \right) \text{ S cm}^{-1} \\ &= 0.01148 \text{ S cm}^{-1} \end{aligned}$$

$$\begin{aligned} \text{Molar conductivity, } \Lambda_m &= \frac{\kappa \times 1000}{c} \text{ cm}^3 \text{ L}^{-1} \\ &= \frac{0.01148 \text{ S cm}^{-1} \times 1000 \text{ cm}^3 \text{ L}^{-1}}{0.05 \text{ mol L}^{-1}} \\ &= 229.6 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned}$$

If we want to calculate the values of different quantities in terms of 'm' instead of 'cm',

$$\begin{aligned} \rho &= \frac{RA}{l} \\ &= \frac{5.55 \times 10^3 \Omega \times 0.785 \times 10^{-4} \text{ m}^2}{0.5 \text{ m}} = 87.135 \times 10^{-2} \Omega \text{ m} \end{aligned}$$

$$\kappa = \frac{1}{\rho} = \frac{100}{87.135} \Omega \text{ m} = 1.148 \text{ S m}^{-1}$$

$$\text{and } \Lambda_m = \frac{\kappa}{c} = \frac{1.148 \text{ S m}^{-1}}{50 \text{ mol m}^{-3}} = 229.6 \times 10^{-4} \text{ S m}^2 \text{ mol}^{-1}$$

LECT-6

Variation of Conductivity and Molar Conductivity with Concentration:

Specific conductance is the conductance of 1 cm^3 of solution of an electrolyte. Thus more is the no. of ions per unit volume of the solution; more is the specific conductance / conductivity.

Specific conductance \propto Concentration.

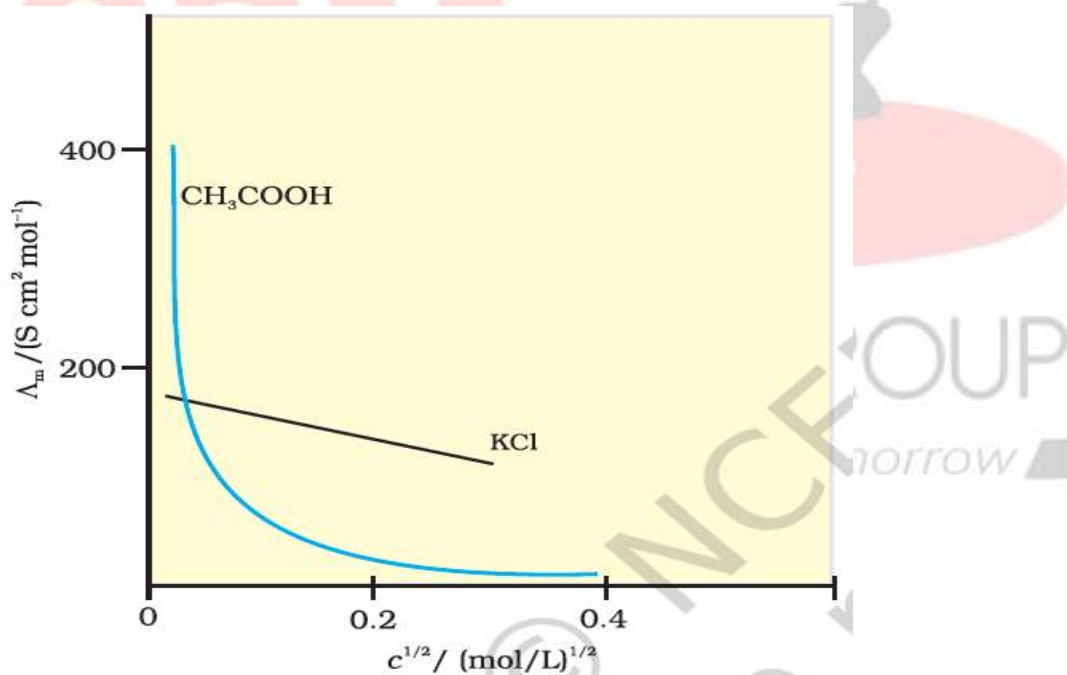
However with dilution, concentration of solution decreases so that the no. of ions per unit volume decreases and hence specific conductance decreases.

Molar conductance (Λ_m)

Molar conductivity (Λ_m) for both the strong and weak electrolyte increases on decreasing the concentration of solution. **In case of strong electrolytes**, molar conductance is slowly increased with decrease in concentration. Because strong electrolytes are almost ionised even at high concentration. But with increase in concentration, molar conductance slowly decreases. According to **Debye-Huckel Onsager** equation.

$$\Lambda_m = \Lambda_m^0 - A\sqrt{C}$$

The graph between Λ_m Vs \sqrt{C} is a straight line with intercept equal to Λ_m^0 and slope equal to $-A$. The value of the constant A for a given solvent and temperature depends on the type of electrolyte. i.e the charges on the cation and anion produced on the dissociation of the electrolyte in the solution. **In case of weak electrolytes** molar conductance increases sharply with dilution. But suddenly decreases with decrease in dilution. Because weak electrolytes are weakly ionized. Graphically it can be represented as-



From the graph it is observed that, indicates molar conductance at zero (0) concentration or at infinite dilution. This is called limiting molar conductance. There is limiting molar conductance for strong electrolyte i.e., molar conductance of strong electrolyte is easily calculated. But it is difficult to calculate for weak electrolytes. Molar conductance of weak electrolytes can be calculated indirectly by using Kohlrausch's law.

If α is the degree of dissociation, then α is equal to

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^0}$$

Here Λ_m^c = Molar conductance at certain concentration.

Λ_m^0 = Molar conductance at zero concentration

But we know that for a weak electrolyte like acetic acid

$$K_a = \frac{c\alpha^2}{(1-\alpha)} = \frac{c\Lambda_m^2}{\Lambda_m^{0^2} \left(1 - \frac{\Lambda_m}{\Lambda_m^0}\right)} = \frac{c\Lambda_m^2}{\Lambda_m^0 (\Lambda_m^0 - \Lambda_m)}$$

Kohlrausch's Law of Independent migration of Ions:-

It states that limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte.

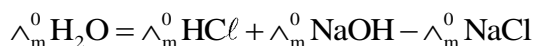
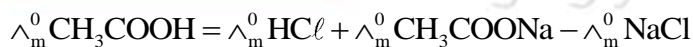
i.e $\Lambda_m^0 \text{NaCl} = \lambda^0 \text{Na}^+ + \lambda^0 \text{Cl}^-$

In general, $\Lambda_m^0 = \nu_+ \lambda_+^0 + \nu_- \lambda_-^0$

Hence λ_+^0 and λ_-^0 are limiting molar conductivity of the cation and anion respectively.

Application of Kohlrausch's law:-

(a) To calculate the limiting molar conductivity (Λ_m) of weak electrolytes.



(b) To calculate the degree of dissociation (α) and dissociation constant (K_a) of an electrostatic.

$$\text{Degree of dissociation } (\alpha) = \frac{\Lambda_m}{\Lambda_m^0}$$

LECT-7**Electrolytic Cell and electrolysis:-**

Electrolysis is a process of decomposition of an electrolyte by passing electric current. Electrolytic cell is a device in which electrical energy is converted into chemical energy by passing electric current. In an electrolytic cell anode is positive and cathode is negative.

Faraday's Laws of Electrolysis:-

1st law:- It states that the mass of the substance that deposited or liberated during electrolysis is directly proportional to the quantity of charge passed through the electrolyte.

$$m \propto Q, \text{ But } Q = It$$

$$\Rightarrow m \propto It$$

$$\Rightarrow m = Zit$$

I – Current in ampere

t- Time in second

Z – Electro chemical equivalent

$$Z = \frac{\text{Equivalent mass}}{\text{Faradays constant}}$$

One Faraday's = charge of 1 mole of electrons

$$6.023 \times 10^{23} \times 1.602 \times 10^{-19} \text{ C}$$

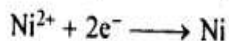
$$= 96487 \text{ C mol}^{-1}$$

3.15 A solution of $\text{Ni}(\text{NO}_3)_2$ is electrolyzed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Sol.

Quantity of electricity passed

$$= (5\text{A}) \times (20 \times 60 \text{ sec.}) = 6000\text{C}$$



Thus, 2F, i.e., $2 \times 96500\text{C}$ of charge deposit

= 1 mole of Ni = 58.7 g

\therefore 6000 C of charge will deposit

$$= \frac{58.7 \times 6000}{2 \times 96500} = 1.825 \text{ g of Ni.}$$

2nd Law:- When two or more electrolytic cells are connected in series and same quantity of charge is passed through them the mass of the substances deposited is directly proportional to their chemical equivalent weight.

$$\text{i.e. } \frac{w_1}{E_1} = \frac{w_2}{E_2} = \frac{w_3}{E_3}$$

w_1, w_2, w_3 are masses and E_1, E_2, E_3 are equivalent masses of substances 1, 2 and 3 respectively.

EDUCATIONAL GROUP

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3.16 Three electrolytic cells A, B, C containing solutions of ZnSO_4 , AgNO_3 and CuSO_4 , respectively are connected in series. A steady current of 1.5 amperes was passed through them until 45 g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

Sol.

Given: $I = 1.5 \text{ A}$, $W = 1.45 \text{ g of Ag}$, $t = ?$, $E = 108$, $n = 1$

Using Faraday's 1st law of electrolysis $W = ZIt$

$$\text{or, } W = \frac{E}{nF} It$$

$$\text{or, } t = \frac{1.45 \times 96500}{1.5 \times 108} = 863.73 \text{ seconds.}$$

Now for Cu, $W_1 = 1.45 \text{ g of Ag}$, $E_1 = 108$, $W_2 = ?$, $E_2 = 31.75$

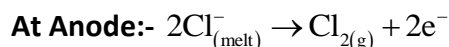
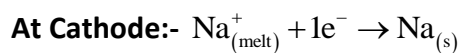
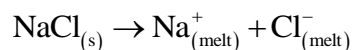
From Faraday's 2nd law of electrolysis $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

$$\frac{1.45}{W_2} = \frac{108}{31.75} \therefore W_2 = \frac{1.45 \times 31.75}{108}$$

Products of electrolysis: - It depends on

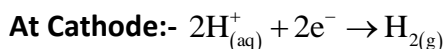
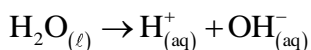
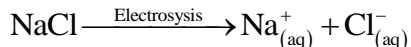
- Nature of material being electrolysed and the type of electrodes being used.
- If the electrode is inert (i.e. platinum or gold or graphite) it does not participate in the chemical reaction and acts only as source for electrons.
- If the electrode is reactive it participates in the electrode reaction. Thus the product of electrolysis may be different for reactive and inert electrodes.
- The products of electrolysis depend on the different oxidising and reducing species present in the electrolytic cell and their standard electrode potentials.

(a) Product of Electrolysis of molten NaCl

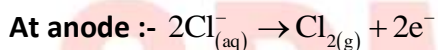


So during electrolysis of molten NaCl, sodium metal will be deposited at cathode and chlorine gas will be liberated at anode.

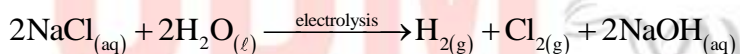
(b) Product of Electrolysis of Aqueous NaCl.



This is because standard reduction potential of $\text{H}^{+} / \text{H}_2$ is greater than $\text{E}^{\ominus} \text{Na}^{+} / \text{Na}$

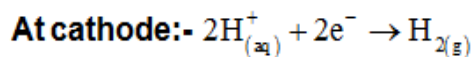
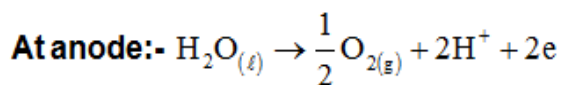
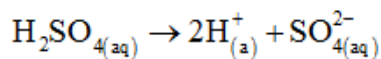


This is because of over potential of oxygen. Therefore,

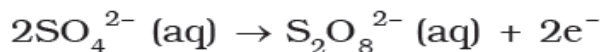


(c) Product of electrolysis of dilute H_2SO_4

During electrolysis of dilute H_2SO_4 the product are $\text{H}_{2(\text{g})}$ at cathode and $\text{O}_{2(\text{g})}$ at anode

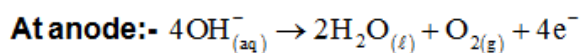
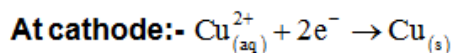
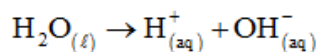
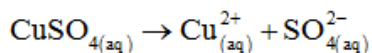


At higher concentrations of H_2SO_4 , oxidation takes place at anode as

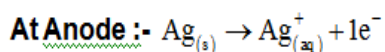
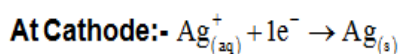


(d) | Product of electrolysis of aqueous copper sulphate using platinum electrode:-

During electrolysis, copper is deposited at cathode and oxygen is liberated at anode

**(e) Product of Electrolysis of AgNO₃ using silver electrodes:-**

During electrolysis impure silver will be oxidized at anode and pure silver will be deposited at cathode.

**BATTERIES**

Electrochemical cells using salt bridge have high internal resistance. As a result we can not draw large current from them and their voltage drops sharply. Thus an electrochemical cell to be used as commercial cell do not drop voltage during use. If a number of cells are connected in series then the arrangement is called battery. The commercial cells are classified as three types such as-

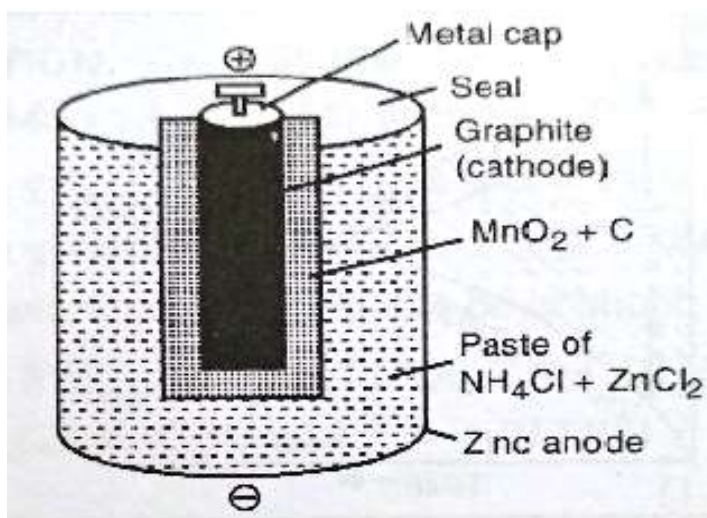
i) Primary cells

ii) Secondary cells

iii) Fuel cells

PRIMARY CELLS:

Primary cells are those in which redox reaction occurs only once and the cell becomes dead after some times and can not be used again. In other words the cell reaction is not reversible. Such primary cells are called dry cells. These cells are used in torches, radio receivers, electronic calculators, hearing aids, etc. Example of Primary cells:- i) Leclanche cell ii) Alkaline battery iii) Mercury battery iv) Nickel-Cadmium(Nicad) battery .



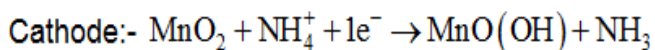
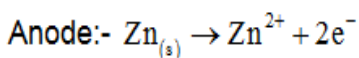
It consists of

Anode – Zinc container

Cathode: - Carbon (graphite) rod surrounded by powdered MnO_2 and carbon

Electrolytic: - Moist paste of NH_4Cl and ZnCl_2

The reaction that takes place is



In the reaction at cathode manganese is reduced from the +4 oxidation state to the +3 oxidation state ammonia produced in the reaction forms complex with Zn^{2+} to give $[\text{Zn}(\text{NH}_3)_4]^{2+}$. The cell potential is nearly 1.5v. It is used in transistor.

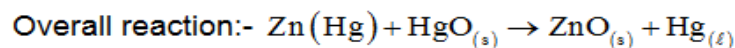
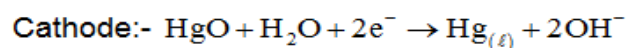
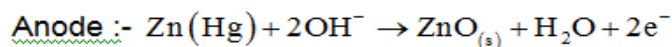
(b) Mercury Cell:- it consists of

Anode:- Zinc – mercury amalgam

Cathode :- Paste of HgO and carbon

Electrolyte paste of KOH and ZnO

The electrode reaction for the cell is



The cell potential is remains constant during its life as the overall reaction does not involve any ion in solution whose concentration changes during its life time. The cell potential is approximately 1.35V. It is used in low current device like hearing aids watches etc.

LECT-8

Secondary Cell/ Battery:- Those type of cell which after use can be recharged by passing current through it in the opposite direction so that it can be used again.

Example:- Lead storage battery.

Nickel – cadmium cell.

(a) Lead storage battery:- It consists of

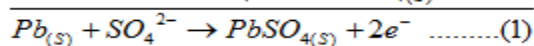
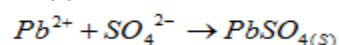
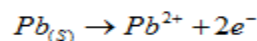
Anode:- Lead

Cathode:- Grid of lead packed with PbO₂

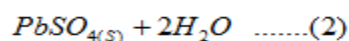
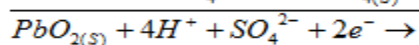
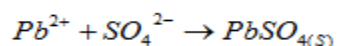
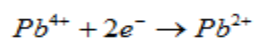
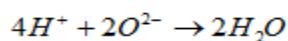
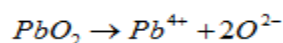
Electrolyte:- 38% solution of H₂SO₄. The cell reaction when the battery is in use.

Discharging Process :

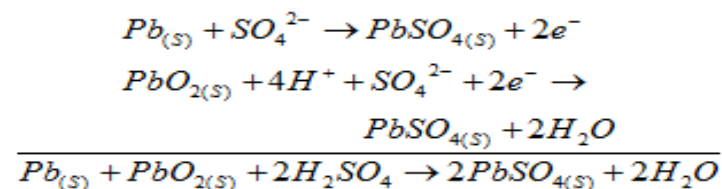
At anode : Pb is oxidised to Pb²⁺ which then react with SO₄²⁻ ions generated from H₂SO₄ to form PbSO₄.



At cathode : The reactions are :

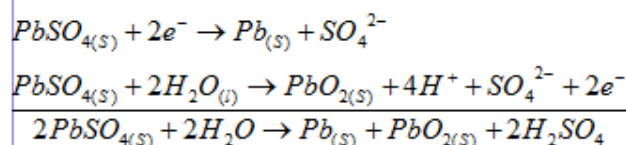


The net cell reaction in the discharging process is obtained by adding equation (1) and (2)



cell potential is approximately 12V. It is used in automobiles and invertors. On charging the battery the reaction is reversed and PbSO₄ on anode and cathode is converted into Pb and PbO₂ respectively.

The net cell reaction during charging is



(b) Nickel Cadmium Cell (NiCad cell):-

It consists of Anode:- Cadmium Cathode:- Metal grid containing Ni, O₂

Electrolyte:- Most KOH/ NaOH. The overall cell reaction that takes place as

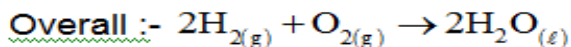
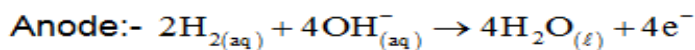
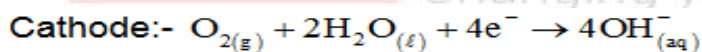


Nickel-cadmium cell has longer life than the lead storage cell but more expensive to manufacture.

Fuel cell:- It is a device in which the energy produced due to combustion of fuel like hydrogen, methane and methanol is converted into electrical energy.

Example:- Hydrogen – oxygen Fuel cell:-

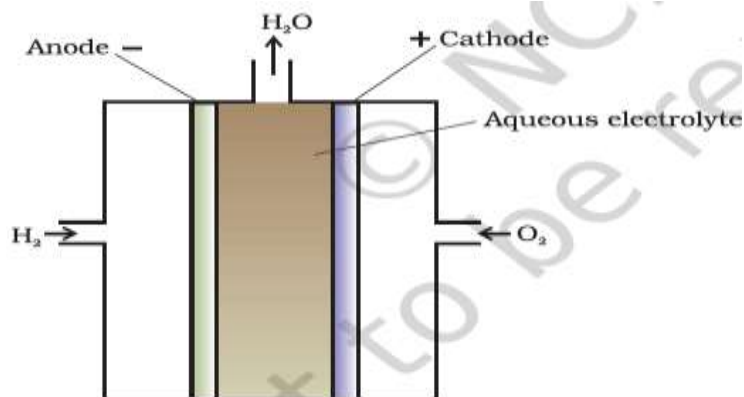
It consists of porous carbon electrodes incorporated with catalyst like finely divided platinum or palladium metal. The electrode reaction that taking place are



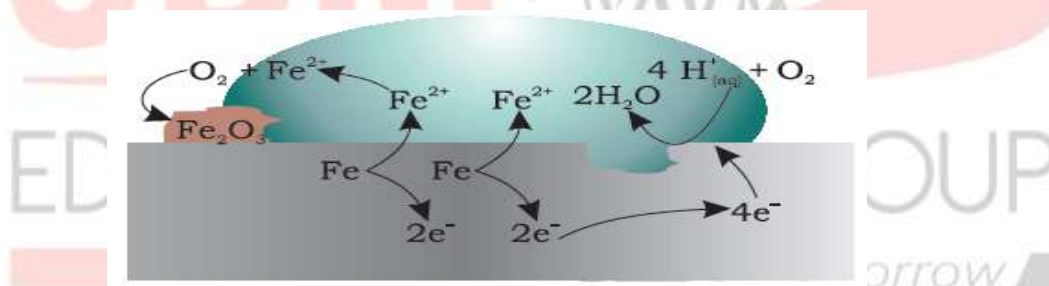
It is used in Apollo space programme.

Advantages of Fuel cell:-

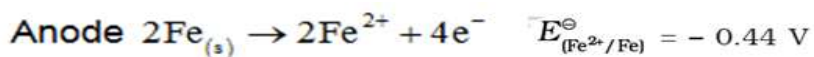
- It has high efficiency of about 70%
- It does not cause pollution
- The by-product water produced can be used for drinking purpose



Corrosion of Metals: The process of slowly eating away of the metal due to attack of the atmospheric gases on the surface of the metal resulting in to the formation of compounds such as oxides, sulphides, carbonates, sulphates etc. is called corrosion. Example: Rusting of iron, Tarnishing of silver, development of green coating on copper and bronze etc. It causes enormous damage to buildings, bridges, ships and to all objects made of metals especially that of iron. In corrosion, a metal is oxidised by loss of electrons to oxygen and formation of oxides. Corrosion of iron (commonly known as rusting) occurs in presence of water and air. Corrosion of iron (rusting) is said to be an electrochemical phenomena as it involves the formation of electrochemical cell on the surface of iron.



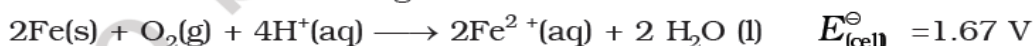
At a particular spot of an object made of iron, oxidation takes place and that spot behaves as anode and we can write the reaction as:



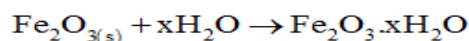
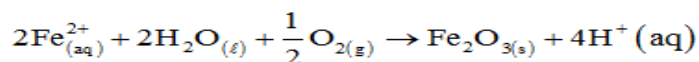
Electrons released at anodic spot move through the metal and go to another spot on the metal and reduce oxygen in presence of H^+ (which is believed to be available from H_2CO_3 formed due to dissolution of carbon dioxide from air into water. Hydrogen ion in water may also be available due to dissolution of other acidic oxides from the atmosphere. This spot behaves as cathode with the reaction:



The overall reaction being:



The ferrous ions are further oxidised by atmospheric oxygen to ferric ions which come out as rust in the form of hydrated ferric oxide as $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ with further production hydrogen ion.



Chemically rust is hydrated ferric oxide.

Prevention of Rusting:-

- Barrier protection:- it can be done by painting oiling, or greasing
- Sacrificial Protection:- it can be done by coating a highly reactive metal on the surface of iron. Example: Galvanization
- Electrical protection (cathodic protection) :- it is used to prevent underground pipes from rusting. Using metal like zinc, aluminum, magnesium
- Using antirust solution:- It is the alkali solution of phosphate and chromate of sodium and potassium. These solutions form a protective insoluble film of iron phosphate which prevents rusting.

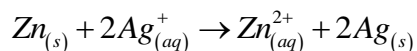
Questions for Electrochemistry:-

Lectures – 1

- Can you store copper sulphate solution in a zinc pot.?
- What would happen, if no salt bridge were used in as electro chemical cell like $\text{Zn}-\text{CuSO}_4$ cell?
- Why does a Galvanic Cell become dead after some time?
- Give the standard electrode potentials $K^+ / K = -2.93\text{v}$, $\text{Ag}^+ / \text{Ag} = 0.80\text{v}$

$\text{Mg}^{2+} / \text{Mg} = -2.37\text{v}$, $\text{Hg}^{2+} / \text{Hg} = 0.79\text{v}$ $\text{Cr}^{3+} / \text{Cr} = -0.74\text{v}$. Arrange these methods in their increasing order of reducing power.

- Depict the Galvanic cell in which the following reaction takes place. Further show



- Which of the electrode is negatively charged?

(b) The carriers of the current in the cell

(c) Individual reaction at each electrolysis.

06. Depict the Galvanic Cell in which the case reaction is $Cu + 2Ag^+ \rightarrow 2Ag + Cu^{2+}$

Lecture – 2

01. Why it is not possible to measure the single electrode potential?

02. How would you determine the standard electrode potential of the system Mg^{2+} / Mg ?

03. Consider a cell given below

$Cu / Cu^{2+} || Cl^- | Cl_2, Pt$ Write the cell reaction that occurs at anode and cathode.

04. Define electrochemical series.

05. What does the negative value of E_{cell}^{θ} indicate?

Lecture – 3

01. Write Nernst equation for single electrode potential $M^{n+} + ne^{-1} \rightarrow M$

02. Write the relationship between E_{cell}^{θ} and Gibb's energy of reaction.

03. Write the relationship between Gibb's energy change of reaction and equilibrium constant.

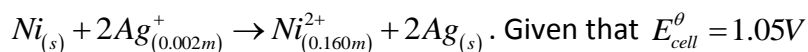
04. Calculate the equilibrium constant of the reaction. $Cu_{(s)} + 2Ag_{(aq)}^+ \rightarrow Cu_{(aq)}^{2+} + 2Ag_{(s)}$

$$E_{cell}^{\theta} = 0.46V$$

05. Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

Lecture – 4

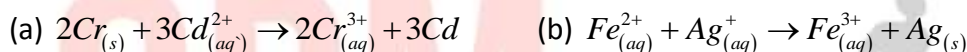
01. Calculate the emf of the cell in which the following reaction takes place.



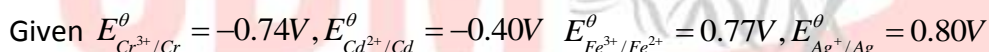
02. The standard electrode potential for Daniell cell is 1.1v. Calculate the standard Gibb's energy for the reaction. $Zn_{(s)} + Cu_{(aq)}^{2+} \rightarrow Zn_{(aq)}^{2+} + Cu_{(s)}$.

03. In the button cells widely used in watches and other devices the following reaction takes place. $Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \rightarrow Zn_{(aq)}^{2+} + 2Ag_{(s)} + ZOH_{(aq)}$.

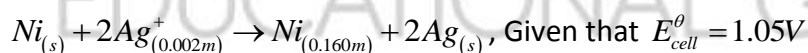
04. Calculate the standard cell potentials of galvanic cell in which the following reaction takes place.



Calculate the $\Delta_r G^{\theta}$ and equilibrium constant for the reaction.



05. Calculate the emf of the cell in which the following reaction takes place.



Lecture – 5

01. Write the relationship between conductivity (K) conductance (G) and Cell constant (G^+)

02. Define cell constant and write the unit

03. State Kohlrausch's law of independent migration of ions.

04. Suggest a way to calculate the \wedge_m^{θ} of water

05. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

06. The conductivity of 0.20 m solution of KCl at 295 K is 0.0248 S cm^{-1} . Calculate its molar conductivity.

07. Calculate the Λ_m^0 for CaCl_2 and MgSO_4 from the following data

$$\lambda^0 \text{Ca}^{2+} = 119.0 \text{ S cm}^2 \text{ mol}^{-1}, \lambda^0 \text{Cl}^- = 76.3 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\lambda^0 \text{Mg}^{2+} = 106.0 \text{ S cm}^2 \text{ mol}^{-1}, \lambda^0 \text{SO}_4^{2-} = 160.0 \text{ S cm}^2 \text{ mol}^{-1}$$

08. Λ_m^0 for NaCl, HCl and NaAc are 126.4, 425.9 and $91.0 \text{ S cm}^2 \text{ mol}^{-1}$ respectively. Calculate Λ_m^0 for HAc.

09. Why does conductivity of a solution decreases with dilution?

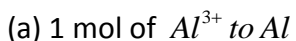
10. The molar conductivity of 0.025 mol L^{-1} methanoic acid (HCOOH) is $46.1 \text{ S cm}^2 \text{ mol}^{-1}$. Calculate its degree of dissociation and dissociation constant.

$$\text{Given } \lambda^0 \text{H}^+ = 349.6 \text{ S cm}^2 \text{ mol}^{-1}, \lambda^0 \text{HCOO}^- = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$$

Lecture – 6

01. State Faraday's 1st law of electrolysis.

02. How much charge is required for the following reduction?



03. How much electricity in Faradays is required to produce?



04. If a current of 0.5 ampere flows through a metallic wire for 240 mins, then how many electrons would flow through the wire?

05. Consider the reaction: $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$. What is the quantity of electricity in coulombs needed to reduce 1 mole of $\text{Cr}_2\text{O}_7^{2-}$?

06. A solution of CuSO_4 is electrolysed for 10 minutes with a current of 1.5 ampere. What is the mass of copper deposited at the cathode?
07. Produce the product of electrolyses of
- (a) A dilute solution of H_2SO_4 with platinum electrodes.
 - (b) An aqueous solution of CuCl_2 with platinum electrodes
 - (c) An aqueous solution of AgNO_3 with platinum electrodes.
08. How will the pH of brine (aqueous NaCl solution) be affected when it is electrolyzed?

Lecture – 7

01. What is a primary battery? Write the out takes place at anode and cathode in a dry cell.
02. Why is mercury cell maintaining constant potential during its life time?
03. What are secondary batteries? Write the reaction that takes place during the charging and discharging of lead strong cell.
04. What are the advantages of nickel cadmium cell over lead strong batteries?
05. For the given cells. Lead storage cell, mercury cell, fuel cell, dry cell. Answer the following.
- (a) Which cell is used in hearing aids?
 - (b) Which cell is used in automobiles and invertors?
 - (c) Which cell does not have long life?
06. Write the formula of complex entity formed when Zn^{2+} react with NH_3 in a dry cell.

Lecture – 8

01. What is a Fuel cell? Write the reaction involve in hydrogen oxygen Fuel cell. What are advantages?

02. Suggest two materials other than hydrogen that can be used as fuels in fuel cell.
03. Explain how rusting of iron is envisaged as setting up of an electrostatic cell?
04. Name the metal which can be used for cathode protection of iron.
05. Define Galvanisation? Name metal which is used for sacrificial protection of iron.
06. Why is rusting of iron is faster in saline water than the ordinary water?
07. Why is alkaline medium inhibits the rusting of iron?





