

## Chapter 2

### ELECTROCHEMISTRY

#### Question and answers

#### 2MARKS

**1. Arrange the following metals in the order in which they displace each other from their salts.**

Al, Cu, Fe, Mg and Zn

**Sol:** Mg, Al, Zn, Fe, Cu.

**2. Given the standard electrode potentials,  $K^+/K = -2.93$  V,  $Ag^+/Ag = 0.80$  V,  $Hg^{2+}/Hg = 0.79$  V,  $Mg^{2+}/Mg = -2.37$  V,  $Cr^{3+}/Cr = 0.74$  V.**

**Arrange these metals in their increasing order of reducing power.**

**Sol:** Higher the oxidation potential more easily it is oxidized and hence greater is the reducing power. Thus, increasing order of reducing power will be  $Ag < Hg < Cr < Mg < K$ .

**3. Depict the galvanic cell in which the reaction**

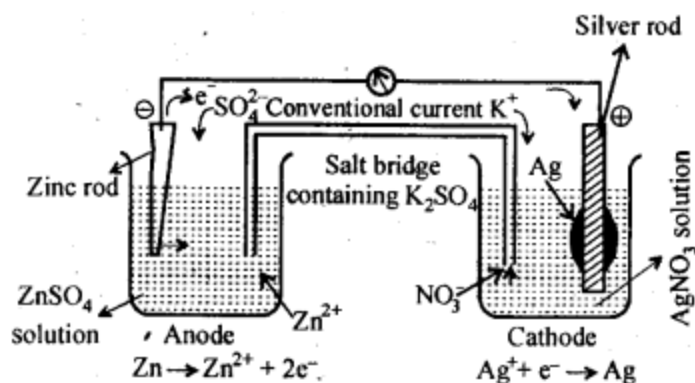
**$Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$  takes place. Further show:**

**(i) Which of the electrode is negatively charged?**

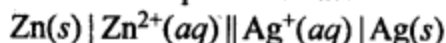
**(ii) The carriers of the current in the cell.**

**(iii) Individual reaction at each electrode.**

**Sol.** The set-up will be similar to as shown below,



The cell will be represented as:



- (i) Anode, i. e, zinc electrode will be negatively charged.
- (ii) The current will flow from silver to copper in the external circuit.
- (iii) At anode:  $\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^{-}$   
At cathode:  $2\text{Ag}^{+}(aq) + 2e^{-} \longrightarrow 2\text{Ag}(s)$

**4. The conductivity of 0.20 M solution of KCl at 298 K is  $0.0248 \text{ S cm}^{-1}$ . Calculate its molar conductivity.**

**Sol:**

$$\Lambda_m = \frac{\kappa \times 1000}{\text{Molarity}} = \frac{0.0248 \text{ S cm}^{-1} \times 1000 \text{ cm}^3 \text{ L}^{-1}}{0.20 \text{ mol L}^{-1}}$$

$$= 124 \text{ S cm}^2 \text{ mol}^{-1}.$$

**5. The resistance of a conductivity cell containing 0.001 M KCl solution at 298 K is  $1500 \Omega$  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:**

$$\text{Cell constant} = \frac{\text{Conductivity}}{\text{Conductance}}$$

$$= \text{Conductivity} \times \text{Resistance}$$

$$= 0.146 \times 10^{-3} \text{ S cm}^{-1} \times 1500 \Omega$$

$$= 0.219 \text{ cm}^{-1}.$$

**6. How much charge is required for the following reductions:**

**(i) 1 mol of  $\text{Al}^{3+}$  to Al?**

**(ii) 1 mol of  $\text{Cu}^{2+}$  to Cu ?**

**(iii) 1 mol of  $\text{MnO}_4^-$  to  $\text{Mn}^{2+}$ ?**

**Sol:** (i) The electrode reaction is  $\text{Al}^{3+} + 3\text{e}^- \longrightarrow \text{Al}$

$\therefore$  Quantity of charge required for reduction of 1 mol of  $\text{Al}^{3+} = 3F = 3 \times 96500 \text{ C} = 289500 \text{ C}$ .

(ii) The electrode reaction is  $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$

$\therefore$  Quantity of charge required for reduction of 1 mol of  $\text{Cu}^{2+} = 2F = 2 \times 96500 = 193000 \text{ C}$ .

(iii) The electrode reaction is  $\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$ .

i.e.,  $\text{Mn}^{7+} + 5\text{e}^- \longrightarrow \text{Mn}^{2+}$ .

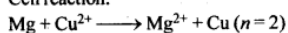
$\therefore$  Quantity of charge required =  $5F$   
 $= 5 \times 96500 \text{ C} = 482500$ .

## 5MARKS

1. Write the Nernst equation and emf of the following cells at 298 K:

- (i)  $\text{Mg}(s) | \text{Mg}^{2+} (0.001 \text{ M}) || \text{Cu}^{2+} (0.0001 \text{ M}) | \text{Cu}(s)$   
 (ii)  $\text{Fe}(s) | \text{Fe}^{2+} (0.001 \text{ M}) || \text{H}^+ (1 \text{ M}) | \text{H}_2(g) (1 \text{ bar}) | \text{Pt}(s)$   
 (iii)  $\text{Sn}(s) | \text{Sn}^{2+} (0.050 \text{ M}) || \text{H}^+ (0.020 \text{ M}) | \text{H}_2(g) (1 \text{ bar}) | \text{Pt}(s)$

(i) Cell reaction:

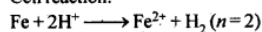


Nernst equation:

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

$$\begin{aligned} \therefore E_{\text{cell}} &= 0.34 - (-2.37) - \frac{0.0591}{2} \log \frac{10^{-3}}{10^{-4}} \\ &= 2.71 - 0.02955 = 2.68 \text{ V.} \end{aligned}$$

(ii) Cell reaction:

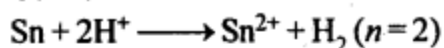


Nernst equation:

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Fe}^{2+}]}{[\text{H}^+]^2}$$

$$\begin{aligned} \therefore E_{\text{cell}} &= 0 - (-0.44) - \frac{0.0591}{2} \log \frac{10^{-3}}{(1)^2} \\ &= 0.44 - \frac{0.0591}{2} \times (-3) \\ &= 0.44 + 0.0887 = 0.5287 \text{ V.} \end{aligned}$$

(iii) Cell reaction:



Nernst equation :

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Sn}^{2+}]}{[\text{H}^+]^2}$$

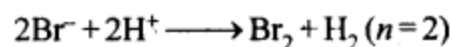
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{0.05}{(0.02)^2}$$

$$= 0 - (-0.14) - \frac{0.0591}{2} \log \frac{0.05}{(0.02)^2}$$

$$= 0.14 - \frac{0.0591}{2} \log 125$$

$$= 0.14 - \frac{0.0591}{2} (2.0969) = 0.078 \text{ V.}$$

(iv) Cell reaction:



(iv)  $\text{Pt}(s) | \text{Br}_2(l) | \text{Br}^- (0.010 \text{ M}) || \text{H}^+ (0.030 \text{ M}) | \text{H}_2(g) (1 \text{ bar}) | \text{Pt}(s)$

Given  $E^\circ_{\text{Mg}^{2+}, \text{Mg}} = -2.37 \text{ V}$ ,

$E^\circ_{\text{Cu}^{2+}, \text{Cu}} = +0.34 \text{ V}$ ,  $E^\circ_{\text{Fe}^{2+}, \text{Fe}} = -0.44 \text{ V}$ .

$E^\circ_{\text{Sn}^{2+}/\text{Sn}} = -0.14 \text{ V}$

$E^\circ_{1/2\text{Br}_2/\text{Br}^-} = +1.08 \text{ V}$

Nernst equation :

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2} \log \frac{1}{[\text{Br}^-]^2 [\text{H}^+]^2}$$

$$= (0 - 1.08) - \frac{0.0591}{2} \log \frac{1}{(0.01)^2 (0.03)^2}$$

$$= -1.08 - \frac{0.0591}{2} \log (1.111 \times 10^7)$$

$$= -1.08 - \frac{0.0591}{2} (7.0457)$$

$$= -1.08 - 0.208 = -1.288 \text{ V}.$$

Thus, oxidation will occur at the hydrogen electrode and reduction will occur on  $\text{Br}_2$  electrode.

2. Conductivity of 0.00241 M acetic acid is  $7.896 \times 10^{-5} \text{ S cm}^{-1}$ . Calculate its molar conductivity. If  $\Lambda_{m0}$ , for acetic acid is  $390.5 \text{ S cm}^2 \text{ mol}^{-1}$ , what is its dissociation constant?

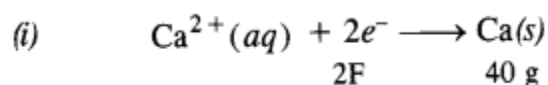
**Sol:**

$$\begin{aligned}\Lambda_m^c &= \frac{\kappa \times 1000}{\text{Molarity}} \\ &= \frac{(7.896 \times 10^{-5} \text{ S cm}^{-1}) \times 1000 \text{ cm}^3 \text{ L}^{-1}}{0.00241 \text{ mol L}^{-1}} \\ &= 32.76 \text{ S cm}^2 \text{ mol}^{-1} \\ \alpha &= \frac{\Lambda_m^c}{\Lambda_m^0} = \frac{32.76}{390.5} = 8.4 \times 10^{-2} \\ K_a &= \frac{C\alpha^2}{1-\alpha} = \frac{0.0024 \times (8.4 \times 10^{-2})^2}{1-0.084} = 1.86 \times 10^{-5}.\end{aligned}$$

3. How much electricity in terms of Faraday is required to produce :

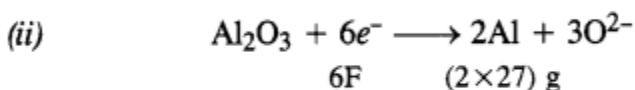
- (i) 20.0 g of Ca from molten  $\text{CaCl}_2$
- (ii) 40.0 g of Al from molten  $\text{Al}_2\text{O}_3$  ?

**Sol:**



To produce 40 g of  $\text{CaCl}_2$ , charge needed = **2F**

To produce 20 g of  $\text{CaCl}_2$ , charge needed = **1F**



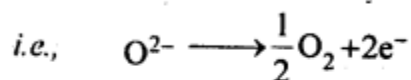
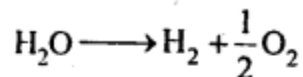
To produce 54 g of Al, charge needed = **6F**

To produce 40 g of Al, charge needed =  $\frac{(40 \text{ g})}{(54 \text{ g})} \times (6F) = \mathbf{4.44F}$ .

4. How much electricity is required in coulomb for the oxidation of (i) 1 mol of  $\text{H}_2\text{O}$  to  $\text{O}_2$  (ii) 1 mol of  $\text{FeO}$  to  $\text{Fe}_2\text{O}_3$

**Sol:**

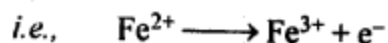
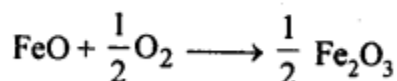
(i) The electrode reaction for 1 mol of  $\text{H}_2\text{O}$  is



$\therefore$  Quantity of electricity required

$$= 2\text{F} = 2 \times 96500 \text{ C} = 193000 \text{ C.}$$

(ii) The electrode reaction for 1 mol of  $\text{FeO}$  is



$\therefore$  Quantity of electricity required =  $1\text{F} = 96500 \text{ C}$

5. 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,  $2\text{F}$ , i.e.,  $2 \times 96500\text{C}$  of charge deposit

$$= 1 \text{ mole of Ni} = 58.7 \text{ g}$$

$\therefore$  6000 C of charge will deposit

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

6. Three electrolytic cells A, B, C containing solutions of  $\text{ZnSO}_4$ ,  $\text{AgNO}_3$  and  $\text{CuSO}_4$ , 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  $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

$$\frac{1.45}{W_2} = \frac{108}{31.75} \therefore W_2 = \frac{1.45 \times 31.75}{108} = 0.426 \text{ g of Cu.}$$

Similarly, for Zn,  $W_1 = 1.45 \text{ g of Ag}$ ,  $E_1 = 108$ ,  $W_2 = ?$ ,  $E_2 = 32.65$

Using formula,  $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

$$\frac{1.45}{W_2} = \frac{108}{32.65}$$

$$\therefore W_2 = \frac{1.45 \times 32.65}{108} = 0.438 \text{ of Zn.}$$



7. Using the standard electrode potentials given in the table, predict if the reaction between the following is feasible.

(a)  $\text{Fe}^{3+}(\text{aq})$  and  $\text{I}^{-}(\text{aq})$

(b)  $\text{Ag}^{+}(\text{aq})$  and  $\text{Cu}(\text{s})$

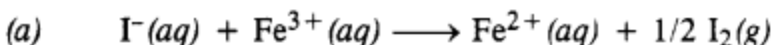
(c)  $\text{Fe}^{3+}(\text{aq})$  and  $\text{Br}^{-}(\text{aq})$

(d)  $\text{Ag}(\text{s})$  and  $\text{Fe}^{3+}(\text{aq})$

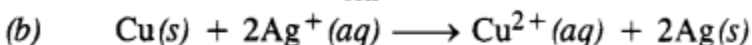
(e)  $\text{Br}_2(\text{aq})$  and  $\text{Fe}^{2+}(\text{aq})$ .

**Sol:**

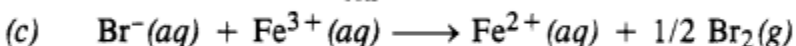
A particular reaction can be feasible if e.m.f. of the cell based on the  $E^{\circ}$  values is positive. Keeping this in mind, let us predict the feasibility of the reactions.



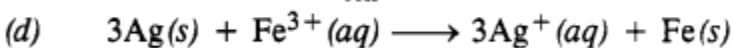
$$E_{\text{cell}}^{\circ} = 0.77 - 0.54 = \mathbf{0.23 \text{ V}} \quad (\text{feasible})$$



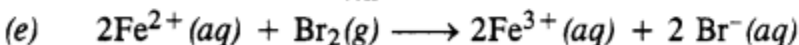
$$E_{\text{cell}}^{\circ} = (0.80 - 0.34) = \mathbf{0.46 \text{ V}} \quad (\text{feasible})$$



$$E_{\text{cell}}^{\circ} = 0.77 - (1.08) = \mathbf{-0.31 \text{ V}} \quad (\text{not feasible})$$



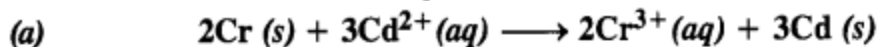
$$E_{\text{cell}}^{\circ} = (0.77 - 0.80) = \mathbf{-0.03 \text{ V}} \quad (\text{not feasible})$$



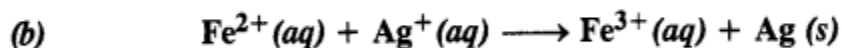
$$E_{\text{cell}}^{\circ} = 1.08 - 0.77 = \mathbf{0.31 \text{ V}} \quad (\text{feasible})$$

## 7MARKS

1. Calculate the standard cell potentials of the galvanic cells in which the following reactions take place.



Given  $E^\circ_{\text{Cr}^{3+}/\text{Cr}} = -0.74 \text{ V}$  ;  $E^\circ_{\text{Cd}^{2+}/\text{Cd}} = -0.40 \text{ V}$



Given  $E^\circ_{\text{Ag}^+/\text{Ag}} = 0.80 \text{ V}$  ;  $E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}} = 0.77 \text{ V}$

Also calculate  $\Delta G^\circ$  and equilibrium constant for the reaction.  
(C.B.S.E. Outside Delhi 2008)

**Sol:**

(a) Calculation of  $E^\circ_{\text{cell}}$ ,

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.40 - (-0.74) = +0.34 \text{ V}$$

Calculation of  $\Delta G^\circ$ ,

$$\begin{aligned}\Delta G^\circ &= -nF E^\circ_{\text{cell}} = -(6 \text{ mol}) \times (96500 \text{ C mol}^{-1}) \times (0.34 \text{ V}) \\ &= -196860 \text{ CV} = -196860 \text{ J} = -196.86 \text{ kJ}\end{aligned}$$

Calculation of Equilibrium Constant ( $K_c$ )

$$\Delta G^\circ = -2.303 RT \log K_c$$

$$\log K_c = \frac{(-)\Delta G^\circ}{2.303 RT} = (-) \frac{(-)196860}{2.303 \times 8.314 \times 298} = 34.501$$

$$K_c = \text{Antilog}(34.501) = 3.17 \times 10^{34}$$

(b) Calculation of  $E^\circ_{\text{cell}}$ ,

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (0.80 - 0.77) = 0.03 \text{ V}$$

Calculation of  $\Delta G^\circ$ ,

$$\begin{aligned}\Delta G^\circ &= -nF E^\circ_{\text{cell}} = -(1 \text{ mol}) \times (96500 \text{ C mol}^{-1}) \times (0.03 \text{ V}) \\ &= -2895 \text{ CV} = -2895 \text{ J} = -2.895 \text{ kJ}\end{aligned}$$

Calculation of Equilibrium Constant ( $K_c$ )

$$\Delta G^\circ = -2.303 RT \log K_c$$

$$\log K_c = (-) \frac{(-\Delta G^\circ)}{2.303 RT} = (-) \frac{(-)2895}{2.303 \times 8.314 \times 298} = 0.5074$$

$$K_c = \text{Antilog}(0.5074) = 3.22$$

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

**Sol:** The reciprocal of resistivity is known as specific conductance or simply conductivity. It is denoted by  $K$  (kappa). Thus, if  $K$  is the specific conductance and  $G$  is the conductance of the solution, then

$$R = \frac{1}{G} \text{ and } \rho = \frac{1}{\kappa} \quad \therefore \frac{1}{G} = \frac{1}{\kappa} \times \frac{l}{A}, \quad \kappa = G \times \frac{l}{A}$$

Now, if  $l = 1 \text{ cm}$  and  $A = 1 \text{ sq.cm}$ , then  $K = G$ .

Hence, conductivity of a solution is defined as the conductance of a solution of 1 cm length and having 1 sq. cm as the area of cross-section. Alternatively, it may be defined as conductance of one centimetre cube of the solution of the electrolyte.

Molar conductivity of a solution at a dilution  $V$  is the conductance of all the ions produced from 1 mole of the electrolyte dissolved in  $V \text{ cm}^3$  of the solution when the electrodes are one cm apart and the area of the electrodes is so large that the whole of the solution is contained between them. It is represented by  $\Delta_m$ .

$$\Delta_m = \frac{\kappa A}{l}$$

Since  $l = 1 \text{ cm}$  and  $A = V$  (volume containing 1 mole of electrolyte)

$$\Delta_m = \kappa V$$

Variation of conductivity and molar conductivity with concentration: Conductivity always decreases with decrease in concentration, for both weak and strong electrolytes. This is because the number of ions per unit volume that carry the current in a solution decreases on dilution.

Molar conductivity increases with decrease in concentration. This is because that total volume,  $V$ , of solution containing one mole of electrolyte also increases. It has been found that decrease in  $K$  on

dilution of a solution is more than compensated by increase in its volume.

3. The conductivity of NaCl at 298 K has been determined at different concentrations and the results are given below:

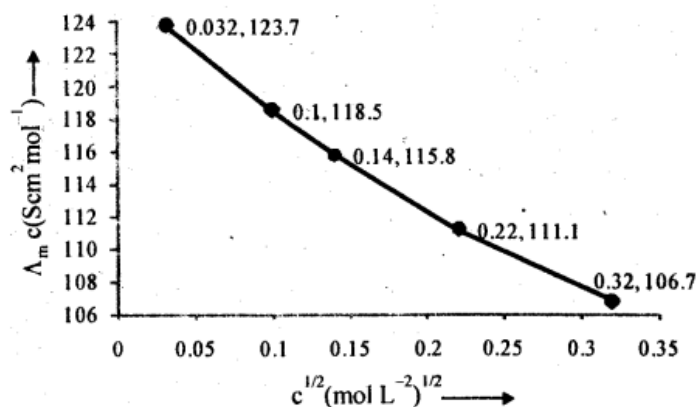
Concentration/M	0.001	0.010	0.020	0.050	0.100
$10^2 \times \kappa/\text{S m}^{-1}$	1.237	11.85	23.15	55.53	106.74

Calculate  $\Lambda_m$  for all concentrations and draw a plot between  $\Lambda_m$  and  $c^{1/2}$ . Find the value of  $\Lambda_m^\circ$ .

**Sol:**

$$\frac{1 \text{ S cm}^{-1}}{100 \text{ S m}^{-1}} = 1 \text{ (unit conversion factor)}$$

Concentration (M)	$\kappa \text{ (S m}^{-1}\text{)}$	$\kappa \text{ (S cm}^{-1}\text{)}$	$\Lambda_m = \frac{1000 \times \kappa}{\text{Molarity}} \text{ (S cm}^2 \text{ mol}^{-1}\text{)}$	$c^{1/2} \text{ (M}^{1/2}\text{)}$
$10^{-3}$	$1.237 \times 10^{-2}$	$1.237 \times 10^{-4}$	$\frac{1000 \times 1.237 \times 10^{-4}}{10^{-3}} = 123.7$	0.0316
$10^{-2}$	$11.85 \times 10^{-2}$	$11.85 \times 10^{-4}$	$\frac{1000 \times 11.85 \times 10^{-4}}{10^{-2}} = 118.5$	0.100
$2 \times 10^{-2}$	$23.15 \times 10^{-2}$	$23.15 \times 10^{-4}$	$\frac{1000 \times 23.15 \times 10^{-4}}{2 \times 10^{-2}} = 115.8$	0.141
$5 \times 10^{-2}$	$55.53 \times 10^{-2}$	$55.53 \times 10^{-4}$	$\frac{1000 \times 55.53 \times 10^{-4}}{5 \times 10^{-2}} = 111.1$	0.224
$10^{-1}$	$106.74 \times 10^{-2}$	$106.74 \times 10^{-4}$	$\frac{1000 \times 106.74 \times 10^{-4}}{10^{-1}} = 106.7$	0.316



$\Lambda^\circ$  = Intercept on  $\Lambda_m$  axis =  $124.0 \text{ S cm}^2 \text{ mol}^{-1}$ , which is obtained by extrapolation to zero concentration.

4. Predict the products of electrolysis in each of the following.

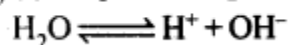
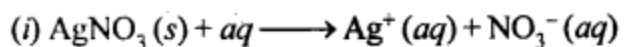
(i) An aqueous solution of  $\text{AgNO}_3$  with silver electrodes.

(ii) An aqueous solution of  $\text{AgNO}_3$  with platinum electrodes.

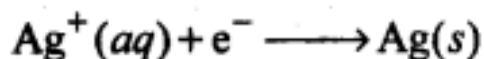
(iii) A dilute solution of  $\text{H}_2\text{SO}_4$  with platinum electrodes.

(iv) An aqueous solution of  $\text{CuCl}_2$  with platinum electrodes.

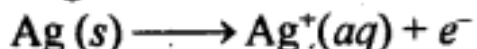
**Sol:**



**At cathode:**  $\text{Ag}^+$  ions have lower discharge potential than  $\text{H}^+$  ions. Hence,  $\text{Ag}^+$  ions will be deposited as Ag in preference to  $\text{H}^+$  ions:

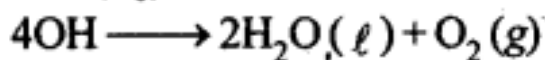
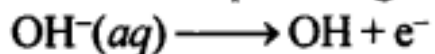


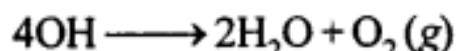
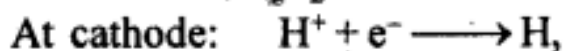
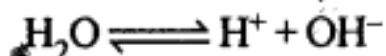
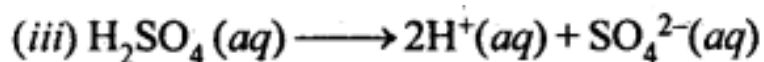
**At anode:** As Ag anode is attacked by  $\text{NO}_3^-$  ions, Ag of the anode will dissolve to form  $\text{Ag}^+$  ions in the solution.



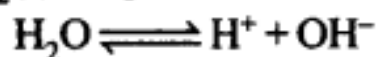
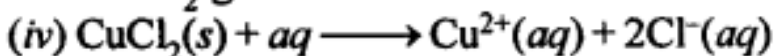
(ii) **At cathode:**  $\text{Ag}^+$  ions have lower discharge potential than  $\text{H}^+$  ions. Hence,  $\text{Ag}^+$  ions will be deposited as Ag in preference to  $\text{H}^+$  ions.

**At anode:** As anode is not attackable, out of  $\text{OH}^-$  and  $\text{NO}_3^-$  ions,  $\text{OH}^-$  ions have lower discharge potential. Hence,  $\text{OH}^-$  ions will be discharged in preference to  $\text{NO}_3^-$  ions, which then decompose to give out  $\text{O}_2$ :

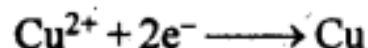




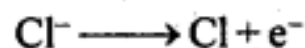
Thus,  $\text{H}_2$  gas is liberated at the cathode and  $\text{O}_2$  gas at the anode.



At cathode:  $\text{Cu}^{2+}$  ions will be reduced in preference to  $\text{H}^+$  ions and copper will be deposited at cathode.



At anode:  $\text{Cl}^-$  ions will be discharged in preference to  $\text{OH}^-$  ions which remains in solution.



Thus, Cu will be deposited on the cathode and  $\text{Cl}_2$  gas will be liberated at the anode.

## SUMMARY

**1. Electrochemistry** is the branch of chemistry which deals with the relationship between electrical energy and chemical energy and inter-conversion of one form into another.

**2. An electrochemical cell** consists of two metallic electrodes dipped in electrolytic solutions. The cells are of two types:

(a) Electrolytic cells (b) Galvanic cells

**3. A galvanic cell** consists of two half cells. Each half cell contains an electrolytic solution and a metallic electrode. The electrode at which- oxidation takes place is called an anode and the electrode at which reduction takes place is called the cathode. The half-cells are separated from each other by means of a porous pot or a salt bridge.

**4. The passage of current from one electrode** to the other indicates the existence of potential difference between them. This difference of potential which causes current to flow from the electrode of higher negative potential is called the electromotive force (emf).

**5. Electrical energy** = Emf (volts) x Quantity of electricity (coulombs)

**6. The potential of SHE** is assigned an arbitrary value of zero.  $E^\circ = 0 \text{ V}$ . It is used as a reference electrode for measuring the standard electrode potentials. .

**7. When the elements** are arranged in order of their standard electrode potentials, a series known as electrochemical series is obtained.

**8. Standard emf of a cell,**

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = E_{\text{Right}}^{\circ} - E_{\text{Left}}^{\circ}$$