

**Question 3.1:**

Arrange the following metals in the order in which they displace each other from the solution of their salts.

Al, Cu, Fe, Mg and Zn

Answer

The following is the order in which the given metals displace each other from the solution of their salts.

Mg, Al, Zn, Fe, Cu

**Question 3.2:**

Given the standard electrode potentials,

$K^+/K = -2.93V$ ,  $Ag^+/Ag = 0.80V$ ,

$Hg^{2+}/Hg = 0.79V$

$Mg^{2+}/Mg = -2.37 V$ ,  $Cr^{3+}/Cr = - 0.74V$

Arrange these metals in their increasing order of reducing power.

Answer

The lower the reduction potential, the higher is the reducing power. The given standard electrode potentials increase in the order of  $K^+/K < Mg^{2+}/Mg < Cr^{3+}/Cr < Hg^{2+}/Hg < Ag^+/Ag$ .

Hence, the reducing power of the given metals increases in the following order:

$Ag < Hg < Cr < Mg < K$

**Question 3.3:**

Depict the galvanic cell in which the reaction  $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$  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.

Answer

The galvanic cell in which the given reaction takes place is depicted as:

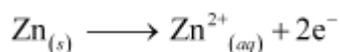


**(i)** Zn electrode (anode) is negatively charged.

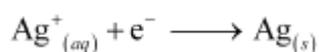


(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.

(iii) The reaction taking place at the anode is given by,

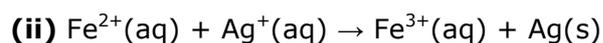
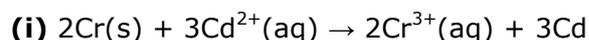


The reaction taking place at the cathode is given by,



#### Question 3.4:

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



Calculate the  $\Delta_r G^\ominus$  and equilibrium constant of the reactions.

Answer

(i)  $E^\ominus_{\text{Cr}^{3+}/\text{Cr}} = 0.74 \text{ V}$

$$E^\ominus_{\text{Cd}^{2+}/\text{Cd}} = -0.40 \text{ V}$$

The galvanic cell of the given reaction is depicted as:



Now, the standard cell potential is

$$\begin{aligned} E^\ominus_{\text{cell}} &= E^\ominus_{\text{R}} - E^\ominus_{\text{L}} \\ &= -0.40 - (-0.74) \\ &= +0.34 \text{ V} \end{aligned}$$

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

In the given equation,

$$n = 6$$

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

$$E^\ominus_{\text{cell}} = +0.34 \text{ V}$$

$$\text{Then, } \Delta_r G^\ominus = -6 \times 96487 \text{ C mol}^{-1} \times 0.34 \text{ V}$$



$$= -196833.48 \text{ CV mol}^{-1}$$

$$= -196833.48 \text{ J mol}^{-1}$$

$$= -196.83 \text{ kJ mol}^{-1}$$

Again,

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

$$\Rightarrow \Delta_r G^\ominus = -2.303 RT \ln K$$

$$\begin{aligned} \Rightarrow \log K &= -\frac{\Delta_r G^\ominus}{2.303 RT} \\ &= \frac{-196.83 \times 10^3}{2.303 \times 8.314 \times 298} \end{aligned}$$

$$= 34.496$$

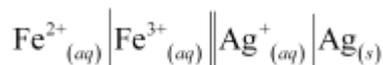
$$\therefore K = \text{antilog}(34.496)$$

$$= 3.13 \times 10^{34}$$

$$\text{(ii)} \quad E^\ominus_{\text{Fe}^{3+}/\text{Fe}^{2+}} = 0.77 \text{ V}$$

$$E^\ominus_{\text{Ag}^+/\text{Ag}} = 0.80 \text{ V}$$

The galvanic cell of the given reaction is depicted as:



Now, the standard cell potential is

$$\begin{aligned} E^\ominus_{\text{cell}} &= E^\ominus_{\text{R}} - E^\ominus_{\text{L}} \\ &= 0.80 - 0.77 \\ &= 0.03 \text{ V} \end{aligned}$$

Here,  $n = 1$ .

$$\text{Then, } \Delta_r G^\ominus = -nFE^\ominus_{\text{cell}}$$

$$= -1 \times 96487 \text{ C mol}^{-1} \times 0.03 \text{ V}$$

$$= -2894.61 \text{ J mol}^{-1}$$

$$= -2.89 \text{ kJ mol}^{-1}$$

$$\text{Again, } \Delta_r G^\ominus = -2.303 RT \ln K$$



$$\Rightarrow \log K = -\frac{\Delta_r G}{2.303 RT}$$
$$= \frac{-2894.61}{2.303 \times 8.314 \times 298}$$

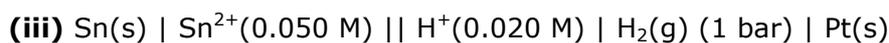
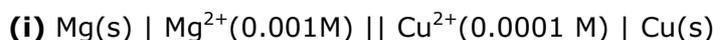
$$= 0.5073$$

$$\therefore K = \text{antilog}(0.5073)$$

$$= 3.2 \text{ (approximately)}$$

**Question 3.5:**

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



Answer

(i) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]}$$
$$= \{0.34 - (-2.36)\} - \frac{0.0591}{2} \log \frac{.001}{.0001}$$
$$= 2.7 - \frac{0.0591}{2} \log 10$$

$$= 2.7 - 0.02955$$

$$= 2.67 \text{ V (approximately)}$$

(ii) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{n} \log \frac{[\text{Fe}^{2+}]}{[\text{H}^+]^2}$$
$$= \{0 - (-0.44)\} - \frac{0.0591}{2} \log \frac{0.001}{1^2}$$
$$= 0.44 - 0.02955(-3)$$

$$= 0.52865 \text{ V}$$

$$= 0.53 \text{ V (approximately)}$$



(iii) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{n} \log \frac{[\text{Sn}^{2+}]}{[\text{H}^+]^2}$$
$$= \{0 - (-0.14)\} - \frac{0.0591}{2} \log \frac{0.050}{(0.020)^2}$$

$$= 0.14 - 0.0295 \times \log 125$$

$$= 0.14 - 0.062$$

$$= 0.078 \text{ V}$$

$$= 0.08 \text{ V (approximately)}$$

(iv) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E_{\text{cell}}^{\ominus} - \frac{0.0591}{n} \log \frac{1}{[\text{Br}^-]^2 [\text{H}^+]^2}$$
$$= (0 - 1.09) - \frac{0.0591}{2} \log \frac{1}{(0.010)^2 (0.030)^2}$$
$$= -1.09 - 0.02955 \times \log \frac{1}{0.00000009}$$
$$= -1.09 - 0.02955 \times \log \frac{1}{9 \times 10^{-8}}$$
$$= -1.09 - 0.02955 \times \log (1.11 \times 10^7)$$
$$= -1.09 - 0.02955 (0.0453 + 7)$$
$$= -1.09 - 0.208$$
$$= -1.298 \text{ V}$$

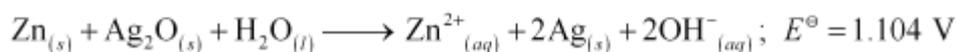
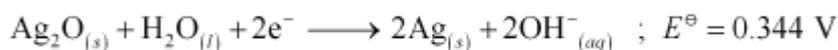
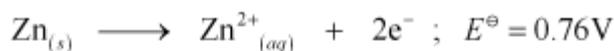
### Question 3.6:

In the button cells widely used in watches and other devices the following reaction takes place:



Determine  $\Delta_r G^\ominus$  and  $E^\ominus$  for the reaction.

Answer



$$\therefore E^{\ominus} = 1.104\text{V}$$

We know that,

$$\Delta_r G^{\ominus} = -nFE^{\ominus}$$

$$= -2 \times 96487 \times 1.04$$

$$= -213043.296\text{J}$$

$$= -213.04\text{kJ}$$

### Question 3.7:

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

Answer

Conductivity of a solution is defined as the conductance of a solution of 1 cm in length and area of cross-section 1 sq. cm. The inverse of resistivity is called conductivity or specific conductance. It is represented by the symbol  $\kappa$ . If  $\rho$  is resistivity, then we can write:

$$\kappa = \frac{1}{\rho}$$

The conductivity of a solution at any given concentration is the conductance ( $G$ ) of one unit volume of solution kept between two platinum electrodes with the unit area of cross-section and at a distance of unit length.

$$G = \kappa \frac{a}{l} = \kappa \cdot 1 = \kappa$$

i.e.,

$$(\text{Since } a = 1, l = 1)$$

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

**Molar conductivity:**



Molar conductivity of a solution at a given concentration is the conductance of volume  $V$  of a solution containing 1 mole of the electrolyte kept between two electrodes with the area of cross-section  $A$  and distance of unit length.

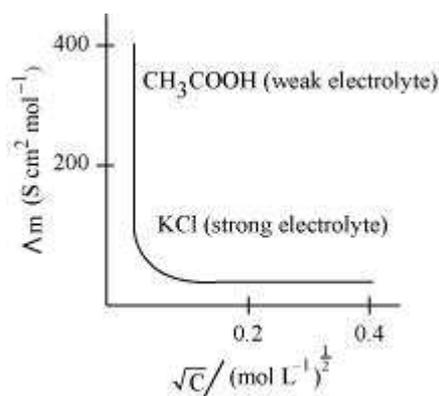
$$\Lambda_m = \kappa \frac{A}{l}$$

Now,  $l = 1$  and  $A = V$  (volume containing 1 mole of the electrolyte).

$$\therefore \Lambda_m = \kappa V$$

Molar conductivity increases with a decrease in concentration. This is because the total volume  $V$  of the solution containing one mole of the electrolyte increases on dilution.

The variation of  $\Lambda_m$  with  $\sqrt{c}$  for strong and weak electrolytes is shown in the following plot:



### Question 3.8:

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

Answer

Given,

$$\kappa = 0.0248 \text{ S cm}^{-1}$$

$$c = 0.20 \text{ M}$$

$$\begin{aligned} \therefore \text{Molar conductivity, } \Lambda_m &= \frac{\kappa \times 1000}{c} \\ &= \frac{0.0248 \times 1000}{0.2} \\ &= 124 \text{ Scm}^2\text{mol}^{-1} \end{aligned}$$

**Question 3.9:**

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

Answer

Given,

$$\text{Conductivity, } \kappa = 0.146 \times 10^{-3} \text{ S cm}^{-1}$$

$$\text{Resistance, } R = 1500 \Omega$$

$$\therefore \text{Cell constant} = \kappa \times R$$

$$= 0.146 \times 10^{-3} \times 1500$$

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

**Question 3.10:**

The conductivity of sodium chloride 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  $\Lambda_m^0$ .

Answer

Given,

$$\kappa = 1.237 \times 10^{-2} \text{ S m}^{-1}, c = 0.001 \text{ M}$$

$$\text{Then, } \kappa = 1.237 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.0316 \text{ M}^{1/2}$$

$$\therefore \Lambda_m = \frac{\kappa}{c}$$
$$= \frac{1.237 \times 10^{-4} \text{ S cm}^{-1}}{0.001 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}}$$

$$= 123.7 \text{ S cm}^2 \text{ mol}^{-1}$$

Given,

$$\kappa = 11.85 \times 10^{-2} \text{ S m}^{-1}, c = 0.010 \text{ M}$$

$$\text{Then, } \kappa = 11.85 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.1 \text{ M}^{1/2}$$



$$\begin{aligned}\therefore \Lambda_m &= \frac{\kappa}{c} \\ &= \frac{11.85 \times 10^{-4} \text{ S cm}^{-1}}{0.010 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} \\ &= 118.5 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Given,

$$\kappa = 23.15 \times 10^{-2} \text{ S m}^{-1}, c = 0.020 \text{ M}$$

$$\text{Then, } \kappa = 23.15 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.1414 \text{ M}^{1/2}$$

$$\begin{aligned}\therefore \Lambda_m &= \frac{\kappa}{c} \\ &= \frac{23.15 \times 10^{-4} \text{ S cm}^{-1}}{0.020 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} \\ &= 115.8 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Given,

$$\kappa = 55.53 \times 10^{-2} \text{ S m}^{-1}, c = 0.050 \text{ M}$$

$$\text{Then, } \kappa = 55.53 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.2236 \text{ M}^{1/2}$$

$$\begin{aligned}\therefore \kappa &= \frac{\kappa}{c} \\ &= \frac{55.53 \times 10^{-4} \text{ S cm}^{-1}}{0.050 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} \\ &= 111.1 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Given,

$$\kappa = 106.74 \times 10^{-2} \text{ S m}^{-1}, c = 0.100 \text{ M}$$

$$\text{Then, } \kappa = 106.74 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.3162 \text{ M}^{1/2}$$

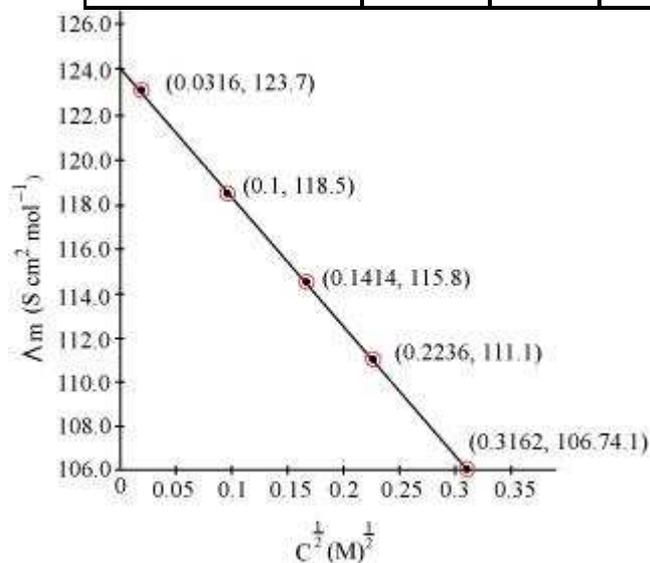
$$\begin{aligned}\therefore \Lambda_m &= \frac{\kappa}{c} \\ &= \frac{106.74 \times 10^{-4} \text{ S cm}^{-1}}{0.100 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} \\ &= 106.74 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Now, we have the following data:

|                            |        |     |        |        |        |
|----------------------------|--------|-----|--------|--------|--------|
| $c^{1/2} / \text{M}^{1/2}$ | 0.0316 | 0.1 | 0.1414 | 0.2236 | 0.3162 |
|----------------------------|--------|-----|--------|--------|--------|



|  |       |       |       |       |        |
|--|-------|-------|-------|-------|--------|
| $\Lambda_m$ (S cm <sup>2</sup> mol <sup>-1</sup> ) | 123.7 | 118.5 | 115.8 | 111.1 | 106.74 |
|--|-------|-------|-------|-------|--------|



Since the line intercepts  $\Lambda_m$  at 124.0 S cm<sup>2</sup> mol<sup>-1</sup>,  $\Lambda_m^0 = 124.0$  S cm<sup>2</sup> mol<sup>-1</sup>.

### Question 3.11:

Conductivity of 0.00241 M acetic acid is  $7.896 \times 10^{-5}$  S cm<sup>-1</sup>. Calculate its molar conductivity and if  $\Lambda_m^0$  for acetic acid is 390.5 S cm<sup>2</sup> mol<sup>-1</sup>, what is its dissociation constant?

Answer

Given,  $\kappa = 7.896 \times 10^{-5}$  S m<sup>-1</sup>

$c = 0.00241$  mol L<sup>-1</sup>

$$\begin{aligned} \text{Then, molar conductivity, } \Lambda_m &= \frac{\kappa}{c} \\ &= \frac{7.896 \times 10^{-5} \text{ S cm}^{-1}}{0.00241 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}} \\ &= 32.76 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned}$$

Again,  $\Lambda_m^0 = 390.5$  S cm<sup>2</sup> mol<sup>-1</sup>

$$\text{Now, } \alpha = \frac{\Lambda_m}{\Lambda_m^0} = \frac{32.76 \text{ S cm}^2 \text{ mol}^{-1}}{390.5 \text{ S cm}^2 \text{ mol}^{-1}}$$



$$= 0.084$$

$$\therefore \text{Dissociation constant, } K_a = \frac{c\alpha^2}{(1-\alpha)}$$
$$= \frac{(0.00241 \text{ mol L}^{-1})(0.084)^2}{(1-0.084)}$$

$$= 1.86 \times 10^{-5} \text{ mol L}^{-1}$$

**Question 3.12:**

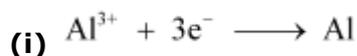
How much charge is required for the following reductions:

(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  $\text{Mn}^{2+}$ .

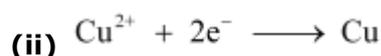
Answer



$\therefore$  Required charge = 3 F

$$= 3 \times 96487 \text{ C}$$

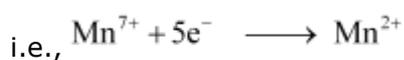
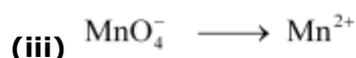
$$= 289461 \text{ C}$$



$\therefore$  Required charge = 2 F

$$= 2 \times 96487 \text{ C}$$

$$= 192974 \text{ C}$$



$\therefore$  Required charge = 5 F

$$= 5 \times 96487 \text{ C}$$

$$= 482435 \text{ C}$$

**Question 3.13:**

How much electricity in terms of Faraday is required to produce

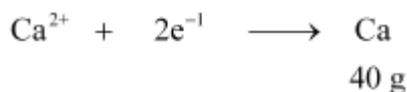
(i) 20.0 g of Ca from molten  $\text{CaCl}_2$ .



(ii) 40.0 g of Al from molten  $\text{Al}_2\text{O}_3$ .

Answer

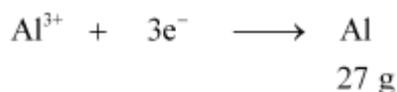
(i) According to the question,



Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium  $= \frac{2 \times 20}{40} \text{ F}$   
= 1 F

(ii) According to the question,



Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al  $= \frac{3 \times 40}{27} \text{ F}$   
= 4.44 F

### Question 3.14:

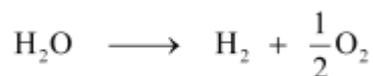
How much electricity is required in coulomb for the oxidation of

(i) 1 mol of  $\text{H}_2\text{O}$  to  $\text{O}_2$ .

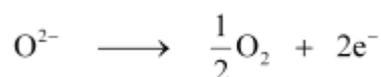
(ii) 1 mol of  $\text{FeO}$  to  $\text{Fe}_2\text{O}_3$ .

Answer

(i) According to the question,



Now, we can write:

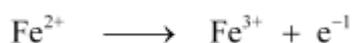


Electricity required for the oxidation of 1 mol of  $\text{H}_2\text{O}$  to  $\text{O}_2$  = 2 F

=  $2 \times 96487 \text{ C}$

= 192974 C

(ii) According to the question,



Electricity required for the oxidation of 1 mol of FeO to Fe<sub>2</sub>O<sub>3</sub> = 1 F  
= 96487 C

**Question 3.15:**

A solution of Ni(NO<sub>3</sub>)<sub>2</sub> is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Answer

Given,

Current = 5A

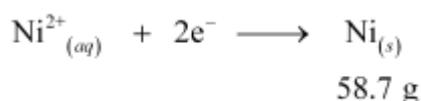
Time = 20 × 60 = 1200 s

∴ Charge = current × time

= 5 × 1200

= 6000 C

According to the reaction,



Nickel deposited by 2 × 96487 C = 58.71 g

Therefore, nickel deposited by 6000 C =  $\frac{58.71 \times 6000}{2 \times 96487}$  g  
= 1.825 g

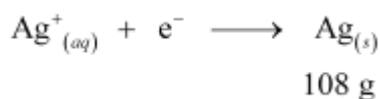
Hence, 1.825 g of nickel will be deposited at the cathode.

**Question 3.16:**

Three electrolytic cells A,B,C containing solutions of ZnSO<sub>4</sub>, AgNO<sub>3</sub> and CuSO<sub>4</sub>, respectively are connected in series. A steady current of 1.5 amperes was passed through them until 1.45 g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

Answer

According to the reaction:



i.e., 108 g of Ag is deposited by 96487 C.

$$\text{Therefore, 1.45 g of Ag is deposited by} = \frac{96487 \times 1.45}{108} \text{ C}$$
$$= 1295.43 \text{ C}$$

Given,

$$\text{Current} = 1.5 \text{ A}$$

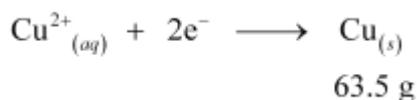
$$\therefore \text{Time} = \frac{1295.43}{1.5} \text{ s}$$

$$= 863.6 \text{ s}$$

$$= 864 \text{ s}$$

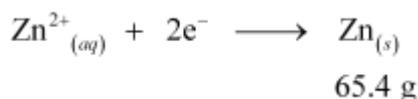
$$= 14.40 \text{ min}$$

Again,



i.e.,  $2 \times 96487$  C of charge deposit = 63.5 g of Cu

$$\text{Therefore, 1295.43 C of charge will deposit} = \frac{63.5 \times 1295.43}{2 \times 96487} \text{ g}$$
$$= 0.426 \text{ g of Cu}$$



i.e.,  $2 \times 96487$  C of charge deposit = 65.4 g of Zn

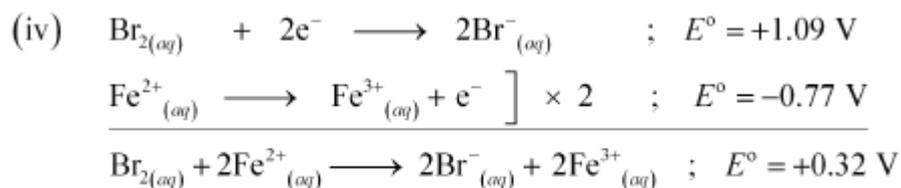
$$\text{Therefore, 1295.43 C of charge will deposit} = \frac{65.4 \times 1295.43}{2 \times 96487} \text{ g}$$
$$= 0.439 \text{ g of Zn}$$

### Question 3.17:

Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:

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





Since  $E^\circ$  for the overall reaction is positive, the reaction between  $\text{Br}_{2(aq)}$  and  $\text{Fe}^{2+}_{(aq)}$  is feasible.

**Question 3.18:**

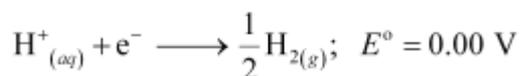
Predict the products of electrolysis in each of the following:

- (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.

Answer

**(i)** At cathode:

The following reduction reactions compete to take place at the cathode.



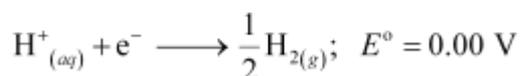
The reaction with a higher value of  $E^\circ$  takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

The Ag anode is attacked by  $\text{NO}_3^-$  ions. Therefore, the silver electrode at the anode dissolves in the solution to form  $\text{Ag}^+$ .

**(ii)** At cathode:

The following reduction reactions compete to take place at the cathode.

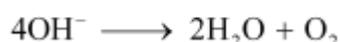




The reaction with a higher value of  $E^\circ$  takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

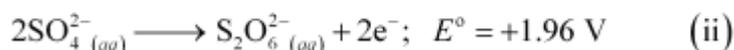
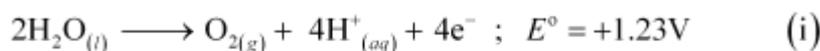
Since Pt electrodes are inert, the anode is not attacked by  $\text{NO}_3^-$  ions. Therefore,  $\text{OH}^-$  or  $\text{NO}_3^-$  ions can be oxidized at the anode. But  $\text{OH}^-$  ions having a lower discharge potential and get preference and decompose to liberate  $\text{O}_2$ .



**(iii)** At the cathode, the following reduction reaction occurs to produce  $\text{H}_2$  gas.



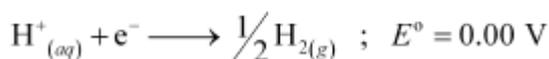
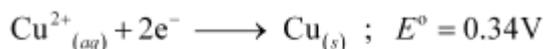
At the anode, the following processes are possible.



For dilute sulphuric acid, reaction (i) is preferred to produce  $\text{O}_2$  gas. But for concentrated sulphuric acid, reaction (ii) occurs.

**(iv)** At cathode:

The following reduction reactions compete to take place at the cathode.

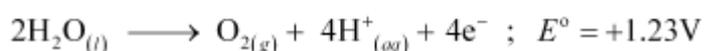
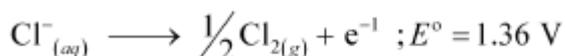


The reaction with a higher value of  $E^\circ$  takes place at the cathode. Therefore, deposition of copper will take place at the cathode.

At

anode:

The following oxidation reactions are possible at the anode.



At the anode, the reaction with a lower value of  $E^\circ$  is preferred. But due to the over-potential of oxygen,  $\text{Cl}^-$  gets oxidized at the anode to produce  $\text{Cl}_2$  gas.

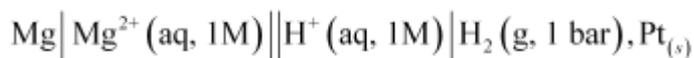
**Text solution****Question 3.1:**

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

Answer

The standard electrode potential of  $\text{Mg}^{2+} | \text{Mg}$  can be measured with respect to the standard hydrogen electrode, represented by  $\text{Pt}_{(s)}, \text{H}_{2(g)} (1 \text{ atm}) | \text{H}^+_{(aq)} (1 \text{ M})$ .

A cell, consisting of  $\text{Mg} | \text{MgSO}_4 (aq, 1 \text{ M})$  as the anode and the standard hydrogen electrode as the cathode, is set up.



Then, the emf of the cell is measured and this measured emf is the standard electrode potential of the magnesium electrode.

$$E^\ominus = E_R^\ominus - E_L^\ominus$$

Here,  $E_R^\ominus$  for the standard hydrogen electrode is zero.



$$\begin{aligned}\therefore E^{\ominus} &= 0 - E_L^{\ominus} \\ &= -E_L^{\ominus}\end{aligned}$$

**Question 3.2:**

Can you store copper sulphate solutions in a zinc pot?

Answer

Zinc is more reactive than copper. Therefore, zinc can displace copper from its salt solution. If copper sulphate solution is stored in a zinc pot, then zinc will displace copper from the copper sulphate solution.



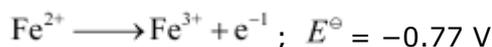
Hence, copper sulphate solution cannot be stored in a zinc pot.

**Question 3.3:**

Consult the table of standard electrode potentials and suggest three substances that can oxidise ferrous ions under suitable conditions.

Answer

Substances that are stronger oxidising agents than ferrous ions can oxidise ferrous ions.



This implies that the substances having higher reduction potentials than +0.77 V can oxidise ferrous ions to ferric ions. Three substances that can do so are  $\text{F}_2$ ,  $\text{Cl}_2$ , and  $\text{O}_2$ .

**Question 3.4:**

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

Answer

For hydrogen electrode,  $\text{H}^{+} + \text{e}^{-} \longrightarrow \frac{1}{2} \text{H}_2$ , it is given that pH = 10

$$\therefore [\text{H}^{+}] = 10^{-10} \text{ M}$$

Now, using Nernst equation:

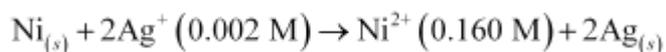
$$E_{\left(\text{H}^{+}/\frac{1}{2}\text{H}_2\right)} = E^{\ominus}_{\left(\text{H}^{+}/\frac{1}{2}\text{H}_2\right)} - \frac{RT}{nF} \ln \frac{1}{[\text{H}^{+}]}$$



$$\begin{aligned} &= E^{\ominus}_{\left(\text{H}^+ / \frac{1}{2}\text{H}_2\right)} - \frac{0.0591}{1} \log \frac{1}{[\text{H}^+]} \\ &= 0 - \frac{0.0591}{1} \log \frac{1}{[10^{-10}]} \\ &= -0.0591 \log 10^{10} \\ &= -0.591 \text{ V} \end{aligned}$$

**Question 3.5:**

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



Given that  $E^{\ominus}_{(\text{cell})} = 1.05 \text{ V}$

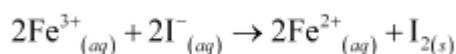
Answer

Applying Nernst equation we have:

$$\begin{aligned} E_{(\text{cell})} &= E^{\ominus}_{(\text{cell})} - \frac{0.0591}{n} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2} \\ &= 1.05 - \frac{0.0591}{2} \log \frac{(0.160)}{(0.002)^2} \\ &= 1.05 - 0.02955 \log \frac{0.16}{0.000004} \\ &= 1.05 - 0.02955 \log 4 \times 10^4 \\ &= 1.05 - 0.02955 (\log 10000 + \log 4) \\ &= 1.05 - 0.02955 (4 + 0.6021) \\ &= 0.914 \text{ V} \end{aligned}$$

**Question 3.6:**

The cell in which the following reactions occurs:





has  $E_{\text{cell}}^{\ominus} = 0.236 \text{ V}$  at 298 K.

Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

Answer

Here,  $n = 2$ ,  $E_{\text{cell}}^{\ominus} = 0.236 \text{ V}$ ,  $T = 298 \text{ K}$

We know that:

$$\begin{aligned}\Delta_r G^{\ominus} &= -nFE_{\text{cell}}^{\ominus} \\ &= -2 \times 96487 \times 0.236 \\ &= -45541.864 \text{ J mol}^{-1} \\ &= -45.54 \text{ kJ mol}^{-1}\end{aligned}$$

Again,  $\Delta_r G^{\ominus} = -2.303RT \log K_c$

$$\begin{aligned}\Rightarrow \log K_c &= -\frac{\Delta_r G^{\ominus}}{2.303 RT} \\ &= -\frac{-45.54 \times 10^3}{2.303 \times 8.314 \times 298}\end{aligned}$$

$$= 7.981$$

$$\therefore K_c = \text{Antilog}(7.981)$$

$$= 9.57 \times 10^7$$

### Question 3.7:

Why does the conductivity of a solution decrease with dilution?

Answer

The conductivity of a solution is the conductance of ions present in a unit volume of the solution. The number of ions (responsible for carrying current) decreases when the solution is diluted. As a result, the conductivity of a solution decreases with dilution.

### Question 3.8:

Suggest a way to determine the  $\Lambda_m^{\ominus}$  value of water.

Answer



Applying Kohlrausch's law of independent migration of ions, the  $\Lambda_m^0$  value of water can be determined as follows:

$$\begin{aligned}\Lambda_{m(\text{H}_2\text{O})}^0 &= \lambda_{\text{H}^+}^0 + \lambda_{\text{OH}^-}^0 \\ &= (\lambda_{\text{H}^+}^0 + \lambda_{\text{Cl}^-}^0) + (\lambda_{\text{Na}^+}^0 + \lambda_{\text{OH}^-}^0) - (\lambda_{\text{Na}^+}^0 + \lambda_{\text{Cl}^-}^0) \\ &= \Lambda_{m(\text{HCl})}^0 + \Lambda_{m(\text{NaOH})}^0 - \Lambda_{m(\text{NaCl})}^0\end{aligned}$$

Hence, by knowing the  $\Lambda_m^0$  values of HCl, NaOH, and NaCl, the  $\Lambda_m^0$  value of water can be determined.

### Question 3.9:

The molar conductivity of 0.025 mol L<sup>-1</sup> methanoic acid is 46.1 S cm<sup>2</sup> mol<sup>-1</sup>.

Calculate its degree of dissociation and dissociation constant. Given  $\lambda^0(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$  and  $\lambda^0(\text{HCOO}^-) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$

Answer

$$C = 0.025 \text{ mol L}^{-1}$$

$$\Lambda_m = 46.1 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\lambda^0(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\lambda^0(\text{HCOO}^-) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\begin{aligned}\Lambda_m^0(\text{HCOOH}) &= \lambda^0(\text{H}^+) + \lambda^0(\text{HCOO}^-) \\ &= 349.6 + 54.6 \\ &= 404.2 \text{ S cm}^2 \text{ mol}^{-1}\end{aligned}$$

Now, degree of dissociation:

$$\begin{aligned}\alpha &= \frac{\Lambda_m(\text{HCOOH})}{\Lambda_m^0(\text{HCOOH})} \\ &= \frac{46.1}{404.2} \\ &= 0.114 \text{ (approximately)}\end{aligned}$$

Thus, dissociation constant:



$$\begin{aligned}K &= \frac{c \alpha^2}{(1-\alpha)} \\ &= \frac{(0.025 \text{ mol L}^{-1})(0.114)^2}{(1-0.114)} \\ &= 3.67 \times 10^{-4} \text{ mol L}^{-1}\end{aligned}$$

**Question 3.10:**

If a current of 0.5 ampere flows through a metallic wire for 2 hours, then how many electrons would flow through the wire?

Answer

$$I = 0.5 \text{ A}$$

$$t = 2 \text{ hours} = 2 \times 60 \times 60 \text{ s} = 7200 \text{ s}$$

$$\text{Thus, } Q = It$$

$$= 0.5 \text{ A} \times 7200 \text{ s}$$

$$= 3600 \text{ C}$$

We know that  $96487 \text{ C} = 6.023 \times 10^{23}$  number of electrons.

Then,

$$\begin{aligned}3600 \text{ C} &= \frac{6.023 \times 10^{23} \times 3600}{96487} \text{ number of electrons} \\ &= 2.25 \times 10^{22} \text{ number of electrons}\end{aligned}$$

Hence,  $2.25 \times 10^{22}$  number of electrons will flow through the wire.

**Question 3.11:**

Suggest a list of metals that are extracted electrolytically.

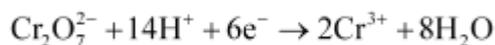
Answer

Metals that are on the top of the reactivity series such as sodium, potassium, calcium, lithium, magnesium, aluminium are extracted electrolytically.

**Question 3.12:**

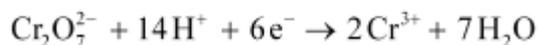
What is the quantity of electricity in coulombs needed to reduce 1 mol of

$\text{Cr}_2\text{O}_7^{2-}$ ? Consider the reaction:



Answer

The given reaction is as follows:



Therefore, to reduce 1 mole of  $\text{Cr}_2\text{O}_7^{2-}$ , the required quantity of electricity will be:

$$= 6 \text{ F}$$

$$= 6 \times 96487 \text{ C}$$

$$= 578922 \text{ C}$$

**Question 3.14:**

Suggest two materials other than hydrogen that can be used as fuels in fuel cells.

Answer

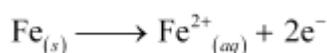
Methane and methanol can be used as fuels in fuel cells.

**Question 3.15:**

Explain how rusting of iron is envisaged as setting up of an electrochemical cell.

Answer

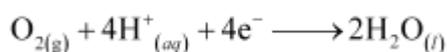
In the process of corrosion, due to the presence of air and moisture, oxidation takes place at a particular spot of an object made of iron. That spot behaves as the anode. The reaction at the anode is given by,



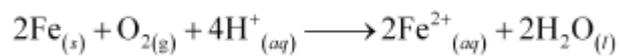
Electrons released at the anodic spot move through the metallic object and go to another spot of the object.

There, in the presence of  $\text{H}^+$  ions, the electrons reduce oxygen. This spot behaves as the cathode. These  $\text{H}^+$  ions come either from  $\text{H}_2\text{CO}_3$ , which are formed due to the dissolution of carbon dioxide from air into water or from the dissolution of other acidic oxides from the atmosphere in water.

The reaction corresponding at the cathode is given by,



The overall reaction is:



Also, ferrous ions are further oxidized by atmospheric oxygen to ferric ions. These ferric ions combine with moisture, present in the surroundings, to form hydrated ferric oxide

$(\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O})$  i.e., rust.

Hence, the rusting of iron is envisaged as the setting up of an electrochemical cell.

$$X_{O_2} = \frac{n_{O_2}}{(n_{O_2} + n_{H_2O})} \approx \frac{n_{O_2}}{n_{H_2O}}$$

$$n_{O_2} = 4.6 \times 10^{-6} \times 55.5 = 2.55 \times 10^{-4} \text{ mol}$$

$$M = 2.55 \times 10^{-4} \text{ M}$$

16. Calculate the mass of compound (molar mass = 256 g/mol) to be dissolved in 75g of benzene to lower its freezing point by 0.48K ( $K_f = 5.12 \text{ K Kg/mol}$ )

**CBSE -**

**2014**

Ans:  $\Delta T_f = K_f \times W_2 \times 1000 / W_1 \times M_2$

$$0.48 = 5.12 \times W_2 \times 1000 / 75 \times 256$$

$$W_2 = 1.8 \text{ g}$$

17. Define an ideal solution and write one of its characteristics

**CBSE -2014**

Ans. **Ideal solution** – The solution which obeys Raoult's law under all conditions of temperature and concentration and during the preparation of which there is no change in enthalpy and volume on mixing the component.

Conditions –

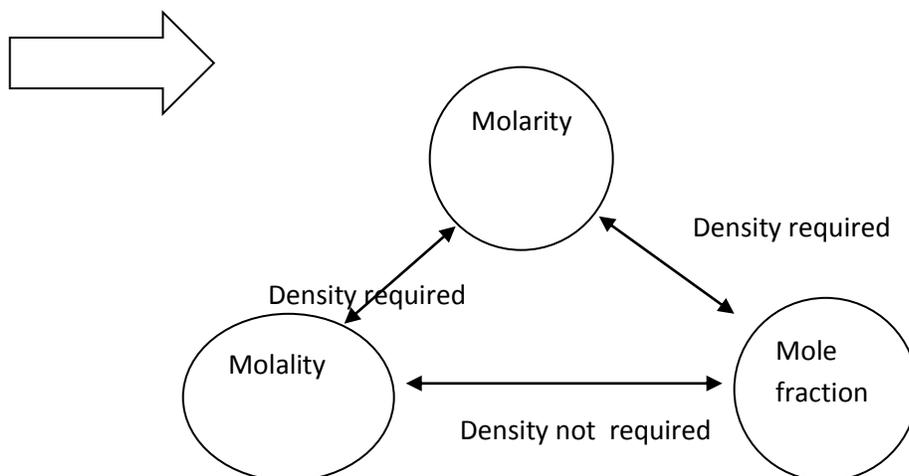
$$P_A = P_A^0 X_A,$$

$$P_B = P_B^0 X_B$$

$$\Delta H_{\text{mix}} = 0,$$

$$\Delta V_{\text{mix}} = 0$$

This is only possible if A-B interaction is same as A-A and B-B interaction nearly ideal solution are –



**Electrochemistry** may be defined as the branch of chemistry which deals with the quantitative study of interrelation ship between chemical energy and electrical energy and inter-conversion of one form into another.

A cell is of two types:-

- I. Galvanic Cell
- II. Electrolytic cell.

In Galvanic cell the chemical energy of a spontaneous redox reaction is converted into electrical work.

In Electrolytic cell electrical energy is used to carry out a non-spontaneous redox reaction.

- The Standard Electrode Potential for any electrode dipped in an appropriate solution is defined with respect to standard electrode potential of hydrogen electrode taken as zero.
- The standard potential of the cell can be obtained by taking the difference of the standard potentials of cathode and anode.

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$$

Concentration dependence of the potentials of the electrodes and the cells are given by Nernst equation.



**Nernst equation can be written as**

$$E_{\text{cell}} = E^0_{\text{cell}} - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

$$nF [A]^a [B]^b$$

At 25<sup>o</sup>C ,  $E_{\text{cell}} = E^0_{\text{cell}} - \frac{0.059}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$

$$n [A]^a [B]^b$$

- The standard potential of the cells are related of standard Gibbs energy.  
 $\Delta_r G^0 = -nFE^0_{\text{cell}}$

- The standard Gibbs energy of the cells is related to equilibrium constant.  
 $\Delta_r G^\circ = -RT \ln K$
- The conductivity,  $K$  of an electrolytic solution depends on the concentration of the electrolyte, nature of solvent and temperature.  
Molar Conductivity,  $\lambda_m$ , is defined as the conductivity power of all the ions produced by one gram mole of an electrolyte in a solution.

$$\lambda_m = \frac{K \times 1000}{\text{Molarity (mol/L)}}$$

|                    | Variation with dilution  |
|--------------------|--|
| Conductivity(k)    | Decreases  |
| Molar conductivity | Increases (Small increase in case of strong electrolytes)<br>(Steep increase in case of weak electrolytes) |

- **Kohlrausch law** of independent migration of ions. The law states that limiting molar conductivity of an electrolyte can be represented as the sum of the individual contribution to the anion and cation of the electrolyte.

#### Faraday's laws of Electrolysis

- The mass of any substance deposited or liberated at any electrode is directly proportional to the quantity of electricity passed through the electrolyte  
 $W = Z I t$

$$\text{Mass(g)} = \frac{\text{Atomic wt.} \times I \text{ (Ampere)} \times t \text{ (s)}}{96500 \times \text{valency}}$$

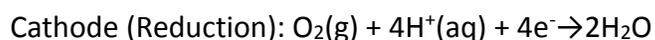
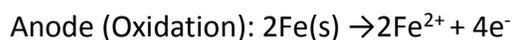
- II. The amount of different substances liberated by the same quantity of electricity passing through the electrolytic solution is proportional to their chemical equivalent weights.

Batteries and full cells are very useful forms of galvanic cells

There are mainly two types of batteries.

- i) Primary batteries eg: Dry cell , Mercury cell
- ii) Secondary batteries(Rechargeable) eg: Lead storage battery ,  
Nickel – cadmium storage cell

- Corrosion of metals is an electrochemical phenomenon.  
In corrosion metal is oxidized by loss of electrons to oxygen and formation of oxides.



Atmospheric Oxidation:



### QUESTION CARRING 1 MARK

1. What is the effect of temperature on molar conductivity?  
Ans. Molar conductivity of an electrolyte increases with increase in temperature.
2. Why is it not possible to measure single electrode potential?  
Ans. It is not possible to measure single electrode potential because the half cell containing single electrode cannot exist independently, as charge cannot flow on its own in a single electrode.
3. Name the factor on which emf of a cell depends:-  
Ans. Emf of a cell depends on following factor-
  - a. Nature of reactants.
  - b. Concentration of solution in two half cells.
  - c. Temperature
  - d. Pressure of gas.
4. What are the units of molar conductivity?  
(  $\text{cm}^2 \text{ohm}^{-1} \text{mol}^{-1}$  or  $\text{Scm}^2\text{mol}^{-1}$  )

5. Write Nernst equation – For the general cell reaction  
**aA+bB→cC+dD**

Ans.  $E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$

6. What is the EMF of the cell when the cell reaction attains equilibrium?

Ans. Zero

7. What is the electrolyte used in a dry cell?

Ans. A paste of  $\text{NH}_4\text{Cl}$ ,  $\text{MnO}_2$  and C

8. How is cell constant calculated from conductance values?

Ans. Cell constant = Resistance x Conductivity

9. What flows in the internal circuit of a galvanic cell.

Ans. Ions

10. Define electrochemical series.

Ans. The arrangement of various electrodes in the decreasing or increasing order of their standard reduction potentials is called electrochemical series.

### QUESTIONS CARRYING TWO MARKS

1. How can you increase the reduction potential of an electrode.? For the reaction



Ans. Nernst equation is:

$$E_{\text{M}^{n+}/\text{M}} = E^{\circ}_{\text{M}^{n+}/\text{M}} - \frac{2.303RT}{nF} \log \frac{1}{[\text{M}^{n+}]}$$

$$nF \quad [M^{n+}]$$

$E_{M^{n+}/M}$  can be increased by

- increase in concentration of  $M^{n+}$  ions in solution
- by decreasing the temperature.

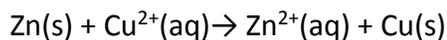
2. Calculate emf of the following cell at 298K



$$\text{Given } E^0 \text{ Zn}^{2+}/\text{Zn} = -0.76 \text{ V}$$

$$E^0 \text{ Cu}^{2+}/\text{Cu} = +0.34 \text{ V}$$

Ans. Cell reaction is as follows.



No of electrons transfers (n)=2

$$T = 298 \text{ K}$$

$$E_{\text{cell}} = (E^0 \text{ Cu}^{2+}/\text{Cu} - E^0 \text{ Zn}^{2+}/\text{Zn}) - \frac{0.059}{n} \log \frac{[\text{Zn}^{2+}(\text{aq})]}{[\text{Cu}^{2+}(\text{aq})]}$$

$$= 0.34 \text{ V} - (-0.76) - \frac{0.059}{2} \log \frac{10^{-4}}{10^{-2}}$$

$$= 0.34 \text{ V} - (-0.76) - 0.02955 \log 10^{-2}$$

$$= 1.10 \text{ V} - 0.02955 \text{ V} \log 10^{-2}$$

$$= 1.10 \text{ V} + 2 \times 0.02955 \text{ V}$$

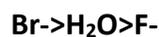
$$= 1.10 \text{ V} + 0.0591 \text{ V}$$

$$= 1.1591 \text{ V}$$

$$= 1.1591 \text{ V}$$

Q 3. Electrolysis of  $\text{KBr(aq)}$  gives  $\text{Br}_2$  at anode but  $\text{KF(aq)}$  does not give  $\text{F}_2$ . Give reason.

Ans. Oxidation takes place at anode. Now higher the oxidation Potential, easier to oxidize. Oxidation potential of  $\text{Br}^-$ ,  $\text{H}_2\text{O}$ , and  $\text{F}^-$  are in the following order.



Therefore in aq. Solution of KBr, Br<sup>-</sup> ions are oxidized to Br<sub>2</sub> in preference to H<sub>2</sub>O. On the other hand, in aq. Solution of KF, H<sub>2</sub>O is oxidized in preference to F<sup>-</sup>. Thus in this case oxidation of H<sub>2</sub>O at anode gives O<sub>2</sub> and no F<sub>2</sub> is produced.

3. What happens when a piece of copper is added to (a) an aq solution of FeSO<sub>4</sub>(b) an Aq solution of FeCl<sub>3</sub>?

a. Nothing will happen when the piece of copper is added to FeSO<sub>4</sub> because reduction potential E<sup>0</sup> Cu<sup>2+</sup>/Cu(0.34) is more than the reduction potential E<sup>0</sup>(Fe<sup>2+</sup>/Fe) (0.44V).

b. Copper will dissolve in an aq solution of FeCl<sub>3</sub> because reduction potential E<sup>0</sup>Fe<sup>3+</sup>/Fe<sup>2+</sup>(0.77V) is more than the reduction potential of E<sup>0</sup>Cu<sup>2+</sup>/Cu(0.34)  
 $\text{Cu(s)} + 2\text{FeCl}_3(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{FeCl}_2(\text{aq})$

4. Define corrosion. Write chemical formula of rust.

Ans. Corrosion is a process of determination of metal as a result of its reaction with air and water, surrounding it. It is due to formulation of sulphides, oxides, carbonates, hydroxides, etc.

Formula of rust- Fe<sub>2</sub>O.XH<sub>2</sub>O

5. What is cell constant?

Ans. Ratio of distance between electrodes and cross section area of electrode

$$G^* = l / A$$

6. Why mercury cell gives the constant voltage.

Ans. Electrolytes are not consuming during the reaction.

7. What is fuel cell, write reaction involved in H<sub>2</sub>-O<sub>2</sub> fuel cell.

#### QUESTION CARRYING THREE MARKS

1. Write any three differences between potential difference and e.m.f.

E.M.F

POTENTIAL DIFFERENCE

1.It is difference between

1.it is difference of potential

electrode potential of two electrodes

between electrode in a closed

when no current is flowing through

circuit.

circuit.

2. it is the maximum voltage obtained

From a cell.

3. it is responsible for steady flow of

Current.

2.it is less than maximum voltage

Obtained from a cell.

3.it is not responsible for steady

Flow of current.

2. Why an electrochemical cell stops working after sometime?

The reduction potential of an electrode depends upon the concentration of solution with which it is in contact.

As the cell works, the concentration of reactants decrease. Then according to Le chatelier's principle it will shift the equilibrium in backward direction. On the other hand if the concentration is more on the reactant side then it will shift the equilibrium in forward direction. When cell works concentration in anodic compartment in cathodic compartment decrease and hence  $E^0$  cathode will decrease. Now EMF of cell is

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$$

A decrease in  $E^0_{\text{cathode}}$  **and a corresponding increase in**  $E^0_{\text{anode}}$  will mean that EMF of the cell will decrease and will ultimately become zero i.e., cell stops working after some time.

3. For the standard cell



$$E^0_{\text{cell}} \text{ } ^{2+}/\text{Cu} = +0.34 \text{ V}$$

$$E^0_{\text{cell}} \text{ } ^{2+} = +0.34 \text{ V}$$

$$E^0 \text{ Ag}^+/\text{Ag} = +0.80 \text{ V}$$

- i. identify the cathode and the anode as the current is drawn from the cell.
- ii. Write the reaction taking place at the electrodes.
- iii. Calculate the standard cell potential.

Ans. 1.From the cell representation

Ag/Ag<sup>+</sup> electrode is cathode and Cu/Cu<sup>+</sup> electrode is anode .



$$\begin{aligned} E^0_{\text{cell}} &= E^0_{\text{cathode}} - E^0_{\text{anode}} \\ &= E^0_{\text{Ag}^+/\text{Ag}} - E^0_{\text{Cu}^{2+}/\text{Cu}} \\ &= +.80 \text{ V} - (+0.34\text{V}) \\ &= +0.80\text{V} - 0.34\text{V} \\ &= 0.46\text{V} \end{aligned}$$

2. Can we store copper sulphate in (i) Zinc vessel (ii) Silver vessel? Give reasons.  
Given E<sup>0</sup> Cu<sup>2+</sup>/Cu = +0.34V, E<sup>0</sup> Zn<sup>2+</sup>/Zn = -0.76V , E<sup>0</sup> Ag<sup>+</sup>/Ag = +0.80V

Ans. A metal having lower reduction potential can displace a metal having higher reduction potential from solution of its salt.

I. Since standard reduction potential of Zn<sup>2+</sup> (E<sup>0</sup> Zn<sup>2+</sup>/Zn = -0.76V) is less than the standard reduction potential of Cu<sup>2+</sup> (E<sup>0</sup> Cu<sup>2+</sup>/Cu = +0.34V), Zn can displace copper from copper sulphate solution but Ag cannot. Thus, CuSO<sub>4</sub> solution can be stored in silver vessel.

3. How many grams of chlorine can be produced by the electrolysis of molten NaCl with a current of 1.02 A for 15 min?

$$\begin{aligned} \text{Mass(g)} &= \frac{\text{Atomic wt.} \times I \text{ (Ampere)} \times t \text{ (s)}}{96500 \times \text{valency}} \\ &= \frac{35.5 \times 1.02 \times 15 \times 60}{96500 \times 1} \\ &= 0.331 \text{ gm} \end{aligned}$$

4. Write the equation showing the effect of concentration on the electrode potential.  
5. Derive the relationship between Gibb's free energy change and the cell potential.

- How Nernst equation can be applied in the calculation of equilibrium constant of any cell reaction.?
- The cell reaction as written is spontaneous if the overall EMF of the cell is positive. Comment on this statement.

### QUESTIONS CARRYING 5 MARKS

- I) Define the following terms **CBSE -2014**
  - Limiting molar conductivity
  - Fuel cell
 II) Resistance of a conductivity cell filled with 0.1 mol/l KCl solution is 100Ω. If the resistance of the same cell when filled with 0.02mol/L KCl solution is 520Ω. Calculate the conductivity and molar conductivity of 0.02mol/L KCl solution. The conductivity of 0.1 mol/L KCl solution is  $1.29 \times 10^{-2} \Omega^{-1} \text{ cm}^{-1}$

**Ans.** NCERT book Page No.

- a) State Faraday's first law of electrolysis. How much charge in terms of Faraday is required for the reduction of 1 mol of Cu<sup>2+</sup> to Cu. **CBSE -2014**
  - Calculate emf of the following cell at 298K. Mg(s)/Mg<sup>2+</sup>(0.1M)//Cu<sup>2+</sup>(0.01)/Cu

**Ans:** a) Amount of substance deposited at an electrode is directly proportional to

the quantity of electricity passed through the electrolyte.

2F

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.059}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]}$$

$$= \frac{2.71}{2} - \frac{0.059}{2} \log \frac{0.1}{0.01} = 2.68\text{V}$$

- Explain the term electrolysis. Discuss briefly the electrolysis of (i) molten NaCl (ii) aqueous sodium chloride solution (iii) molten lead bromide (iv) water.
- state and explain Faraday's laws of electrolysis. What is Electrochemical equivalent?
- What do you understand by 'electrolytic conduction'? what are the factors on which electrolyte conduction depends.? What is the effect of temperature on electrolytic conduction?
- How is electrolytic conductance measured experimentally?
- Describe normal hydrogen electrode and its applications.

### HOT QUESTIONS

1 Mark questions:-

- Why in a concentrated solution, a strong electrolyte shows deviations from Debye-Huckle-Onsagar equation?

Ans:- Because interionic forces of attractions are large.

2. What is the use of Platinum foil in the hydrogen electrode?

A: It is used for inflow and outflow of electrons.

3. Corrosion of motor cars is of greater problem in winter when salts are spread on roads to melt ice and snow. Why?

A: Salts in water are the better electrolytes.

4. Is it safe to stir  $\text{AgNO}_3$  solution with copper spoon? ( $E^0_{\text{Ag}^+/\text{Ag}} = 0.80$  Volt;  $E^0_{\text{Cu}^+/\text{Cu}} = 0.34$  Volt)

Ans: No it is not safe because reacts with  $\text{AgNO}_3$  Solution ( emf will be positive.)

5. Why is it necessary to use salt bridge in A galvanic cell?

Ans: To complete inner circuit and to maintain electrical neutrality of the solution.

2 mark questions:-

1. Why is Li best reducing agent where as Fluorine is best oxidizing agent ?

2. Equilibrium constant is related to  $E^\theta$  cell but not to  $E_{\text{cell}}$ . Explain.

3. Why sodium metal is not obtained at cathode when aq NaCl is electrolysed with Pt electrodes but obtained when molten NaCl is electrolysed ? 2

4. Zn rod weighing 25 g was kept in 100 mL of 1M copper sulphate solution. After

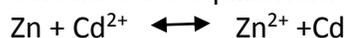
certain time interval, the molarity of  $\text{Cu}^{2+}$  was found to be 0.8 M. What is the

molarity of  $\text{SO}_4^{-2}$  in the resulting solution and what should be the mass of Zn rod after

cleaning and drying ?

8. Which will have greater molar conductivity and why? Sol A. 1mol KCl dissolved in 200cc of the solution or Sol B. 1 mol KCl dissolved in 500cc of the solution.

9. Calculate the equilibrium constant for the reaction



If  $E^0_{\text{Cd}^{2+}/\text{Cd}} = -.403$  v  $E^0_{\text{Zn}^{2+}/\text{Zn}} = -0.763$  v

3/ 5 mark questions:-

1. What do you mean by ( i) negative standard electrode potential and

(ii) positive standard electrode potential ?

2. Which cell is generally used in hearing aids? Name the material of the anode, cathode and the electrolyte. Write the reactions involved.

3. Iron does not rust even if Zinc coating is broken in galvanized iron pipe but rusting occurs much faster if tin coating over iron is broken. Explain.

4. 'Corrosion is an electrochemical phenomenon', explain.

5. Calculate the pH of following cell: Pt, H<sub>2</sub>/ H<sub>2</sub>SO<sub>4</sub>, if its electrode potential is 0.03V.

6. A cell contains two hydrogen electrodes. The negative electrode is in contact with a solution of 10<sup>-5</sup> M H<sup>+</sup> ions. The emf of the cell is 0.118 V at 298 K. Calculate the concentration of the H<sup>+</sup> ions at the positive electrode.

7. Crude copper containing Fe and Ag as contaminations was subjected to electro refining by using a current of 175 A for 6.434 min. The mass of anode was found to decrease by 22.260 g, while that of cathode was increased by 22.011 g. Estimate the % of copper, iron and silver in crude copper.

8. Zinc electrode is constituted at 298 K by placing Zn rod in 0.1 M aq solution of zinc sulphate which is 95 % dissociated at this concentration. What will be the

electrode potential of the electrode given that  $E^\ominus_{Zn^{2+}/Zn} = -0.76$  V. 3

9. At what pH will hydrogen electrode at 298 K show an electrode potential of -0.118 V, when Hydrogen gas is bubbled at 1 atm pressure ? 3

10. Electrolysis of the solution of MnSO<sub>4</sub> in aq sulphuric acid is a method for the preparation of MnO<sub>2</sub> as per the chemical reaction



Passing a current of 27 A for 24 Hrs gives 1 kg of MnO<sub>2</sub>. What is the current efficiency ? What are the reactions occurring at anode and cathode ?

11. What do you mean by Kohlrausch's law: from the following molar conductivities at infinite dilution

$$\lambda^{\circ}_m \text{Ba(OH)}_2 = 457.6 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\lambda^{\circ}_m \text{BaCl}_2 = 240.6 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

$$\lambda^{\circ}_m \text{NH}_4\text{Cl} = 129.8 \Omega^{-1} \text{cm}^2 \text{mol}^{-1}$$

Calculate  $\lambda^{\circ m}$  for  $\text{NH}_4\text{OH}$

**CHAPTER-4 CHEMICAL KINETICS**

**5 Marks Chapter**

**Key points**