

# NCERT Solutions for Class 12 Chemistry

## Chapter 3 Electrochemistry

Class 12 Chemistry chapter 3 Electrochemistry explains different types of cells and the differences between the two. Important concepts such as Gibb's energy and equilibrium constant are also discussed here.

### NCERT IN-TEXT QUESTIONS

Question 1.

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

Answer:

In order to determine  $E^\circ$  value of  $Mg^{2+}/Mg$  electrode, an electrochemical cell is set up in which a Mg electrode dipped in 1 M  $MgSO_4$  solution acts as one half cell (oxidation half cell) while the standard hydrogen electrode acts as the other half cell (reduction half cell). The deflection of voltmeter placed in the cell circuit is towards the Mg- electrode indicating the flow of current. The cell may be represented as :

$Mg/Mg^{2+} (1 M) || H^+(1 M)/H_2(1 atm), Pt$

The reading as given by voltmeter gives  $E^\circ_{cell}$

$$E^\circ_{cell} = E^\circ_{H^+/1/2H_2} - E^\circ_{Mg^{2+}/Mg} = 0 - E^\circ_{Mg^{2+}/Mg}$$

The expected value of standard electrode potential ( $E^\circ$ ) = -2.36 V.

Question 2.

Can you store copper sulphate solution in a zinc pot ?

Answer:

No, it is not possible. The  $E^\circ$  values of the copper and zinc electrodes are as follows :

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$  ;  $E^\circ = -0.76 V$

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$  ;  $E^\circ = +0.34 V$

This shows that zinc is a stronger reducing agent than copper. It will lose electrons to  $Cu^{2+}$  ions and redox reaction will immediately set in.

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Thus, copper sulphate solution cannot be stored in zinc pot.

Question 3.

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

Answer:

Oxidation of  $Fe^{2+}$  converts it to  $Fe^{3+}$ , i.e.,  $Fe^{2+} \rightarrow Fe^{3+} + e^-$  ;  $E^\circ_{ox} = -0.77 V$  Only those substances can oxidise  $Fe^{2+}$  to  $Fe^{3+}$  which are stronger oxidizing agents and have positive reduction potentials greater than 0.77 V, so that EMF of the cell reaction is positive. This is so for elements lying below  $Fe^{3+}/Fe^{2+}$  in the series ex:  $Br_2$ ,  $Cl_2$  and  $F_2$ .

Question 4.

Calculate the potential of hydrogen electrode in contact with a solution with pH equal to 10.

Answer:

For hydrogen electrode,  $H^+ + e^- \rightarrow 1/2H_2$

Applying Nernst equation,

$$\begin{aligned}
 E_{(H^+ / 1/2 H_2)} &= E_{(H^+ / 1/2 H_2)}^\circ - \frac{0.0591}{n} \log \frac{1}{[H^+]} \\
 &= 0 - \frac{0.0591}{1} \log \frac{1}{(10^{-10})} \qquad [pH = 10 ; [H^+] = 10^{-10} M] \\
 &= 0 - \frac{0.0591}{1} \times (10 \log 10) = -0.591 V
 \end{aligned}$$

Question 5.

Calculate e.m.f. of the cell in which the following reaction takes place

$Ni(s) + 2Ag^+(0.002M) \rightarrow Ni^{2+}(0.160M) + 2Ag(s)$  Given that  $E^\circ_{cell} = 1.05 V$ . (C.B.S.E. Outside Delhi 2015)

Answer:

**Cell reaction :**  $Ni(s) + 2Ag^+(0.1 M) \rightarrow Ni^{2+}(0.1 M) + 2 Ag(s)$

$$\begin{aligned}
 E_{cell} &= E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Ni^{2+}]}{[Ag^+]^2} = [0.8 - (-2.25)] - \frac{0.0591}{2} \log \frac{0.1}{(0.1)^2} \\
 &= 1.05 - 0.02955 \log 10 = 1.05 - 0.02955 = 1.02045 V.
 \end{aligned}$$

Question 6.

The cell in which the following reaction occurs

$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s)$  has  $E^\circ_{cell} = 0.236 V$  at 298 K.

Calculate standard Gibbs energy and equilibrium constant for the reaction.

Answer:

The two half reactions are :

$2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$  and  $2I^- \rightarrow I_2 + 2e^-$

For the above reaction,  $n = 2$

$$\begin{aligned}
 \Delta G^\circ &= -nFE^\circ_{cell} = (-2 \text{ mol}) \times (96500 \text{ C mol}^{-1}) \times (0.236 \text{ V}) \\
 &= -455480 \text{ CV} = -455480 \text{ J} = -45.55 \text{ kJ}
 \end{aligned}$$

$$\begin{aligned}
 \text{Now, } \log K_c &= -\frac{\Delta G^\circ}{2.303 RT} = -\frac{(-45.55 \text{ kJ})}{2.303 \times (8.314 \times 10^{-3} \text{ kJ K}^{-1}) \times (298 \text{ K})} = 7.983 \\
 K_c &= \text{Antilog}(7.983) = 9.616 \times 10^7
 \end{aligned}$$

Question 7.

Why does the conductivity of a solution decrease with dilution?

Answer:

Conductivity of a solution is the conductance of ions present in a unit volume of the solutions. On dilution, no. of ions per unit volume decreases. Hence, the conductivity decreases.

Question 8.

Suggest a way to determine the  $\Lambda^\circ_m$  for water.

Answer:

Water ( $H_2O$ ) is a weak electrolyte. Its molar conductance at infinite dilution i.e.,  $\Lambda^\circ_m$  can be

determined in terms of  $\Lambda_m^\circ$  for strong electrolytes. This is in accordance with Kohlrausch's Law.

$$\Lambda_{m\text{H}_2\text{O}}^\circ = \Lambda_{m(\text{HCl})}^\circ + \Lambda_{m(\text{NaOH})}^\circ - \Lambda_{m(\text{NaCl})}^\circ$$

Question 9.

The molar conductance of 0.025 mol L<sup>-1</sup> of methanoic acid is 46.15 cm<sup>2</sup> mol<sup>-1</sup>. Calculate its degree of dissociation and dissociation constant. Given  $\lambda^\circ(\text{H}^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$  and  $\lambda^\circ(\text{HCOO}^-) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$ .

Answer:

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

$$\alpha = \frac{\lambda_m}{\lambda_m^0} = \frac{46.1}{404.2} = 0.114 \Rightarrow \alpha = 11.4\%$$

$$\begin{aligned} K_a &= \frac{C\alpha^2}{1-\alpha} = \frac{0.025 \times (0.114)^2}{1-0.114} \\ &= \frac{0.025 \times 0.114 \times 0.114}{0.886} = 3.67 \times 10^{-4} \end{aligned}$$

Question 10.

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

Answer:

Quantity of charge (Q) passed = Current in amperes x Time in seconds = (0.5 A) X (2 x 60 x 60 s) = 3600 As = 3600 C

No. of electrons flowing through the wire by passing a charge of one faraday (96500 C) = 6.022 x 10<sup>23</sup>

No. of electrons flowing through the wire by passing a charge of 3600 C

$$= 6.022 \times 10^{23} \times \frac{(3600 \text{ C})}{(96500 \text{ C})} = 2.246 \times 10^{22}$$

Question 11.

Suggest a list of metals that are extracted electrolytically.

Answer:

Na, Ca, Mg and Al

Question 12.

Consider the reaction :



What is the quantity of electricity in coulombs needed to reduce 1 mole of Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> ions ?

Answer:

The quantity of electricity in coulombs is 6 F or 6 x 96500 C = 5.76 x 10<sup>5</sup> C.

Question 13.

Write the chemistry of recharging the lead storage battery highlighting all the materials that are involved during recharging.

Answer:

Chemical reactions while recharging :

Chemical reactions while recharging :



Electricity is passed through the electrolyte  $\text{PbSO}_4$  which is converted into  $\text{PbO}_2$  and  $\text{Pb}$ .

Recharging is possible in this case because the  $\text{PbSO}_4$  formed during discharging is a solid and sticks to the electrodes. Therefore, it can either take up or give electrons during recharging.

Question 14.

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

Answer:

Methane and Methanol.

Question 15.

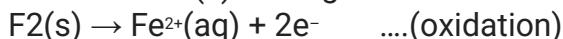
Explain how rusting of iron can be envisaged as the setting up of an electrochemical cell.

Answer:

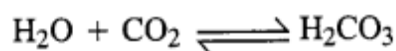
Iron (Fe) is involved in the redox-reaction that is carried in the electrochemical cell which is set up. As a result, it slowly dissolves and the metal surface gets rusted or corroded.

The redox-reaction may be described as follows :

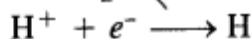
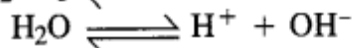
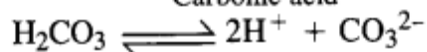
At anode: Fe (s) undergoes oxidation to release electrons



At cathode: The electrons which are released participate in the reduction reaction and combine with  $\text{H}^+$  ions released from carbonic acid ( $\text{H}_2\text{CO}_3$ ) formed by the combination of  $\text{CO}_2$  and  $\text{H}_2\text{O}$  present.



Carbonic acid



...(reduction)

## NCERT EXERCISE

Question 1.

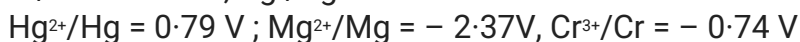
Arrange the following metals in the order in which they displace each other from the solution of their salts: Al, Cu, Fe, Mg and Zn.

Answer:

Mg, Al, Zn, Fe, Cu, Ag.

Question 2.

Given the standard electrode potentials



Arrange these metals in increasing order of their reducing power.

Answer:

Less the electrode potential more will be the reducing power.

$\text{Ag} < \text{Hg} < \text{Cr} < \text{Mg} < \text{K}$ .

Question 3.

Depict the galvanic cell in which the reaction



- Further show :
- (i) which electrode is negatively charged ?
  - (ii) the carriers of the current in the cell.
  - (iii) individual reaction at each electrode. (C.B.S.E. Delhi 2008)

Answer:

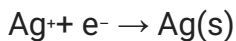
The galvanic cell in which the given reaction takes place is depicted as:



- (i) Zn electrode (anode) is negatively charged
- (ii) Ions are carriers of current in the cell and in the external circuit, current from silver to Zinc.
- (iii) The reaction taking place at the anode is given by,

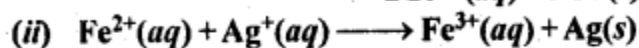
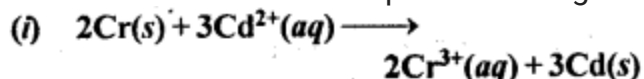


The reaction taking place at the cathode is given



Question 4.

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



Given  $E^\circ_{\text{Cr}^{3+}, \text{Cr}} = -0.74 \text{ V}$ ,

$$E^\circ_{\text{Cd}^{2+}, \text{Cd}} = -0.40 \text{ V}$$

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

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

Answer:

$$\begin{aligned}
 (i) \quad E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\
 &= -0.40 \text{ V} - (-0.74 \text{ V}) = +0.34 \text{ V} \\
 \Delta_r G^{\circ} &= -nFE_{\text{cell}}^{\circ} \\
 &= -6 \times 96500 \text{ C mol}^{-1} \times 0.34 \text{ V} \\
 &= -196860 \text{ C V mol}^{-1} \\
 &= -196860 \text{ J mol}^{-1} = -196.86 \text{ kJ mol}^{-1}
 \end{aligned}$$

$$\begin{aligned}
 -\Delta_r G^{\circ} &= 2.303 RT \log K \\
 196860 &= 2.303 \times 8.314 \times 298 \log K
 \end{aligned}$$

$$\text{or } \log K = 34.5014$$

$$K = \text{Antilog } 34.5014 = 3.172 \times 10^{34}$$

$$(ii) \quad E_{\text{cell}}^{\circ} = +0.80 \text{ V} - 0.77 \text{ V} = +0.03 \text{ V}$$

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

$$\begin{aligned}
 \Delta_r G^{\circ} &= -2.303 RT \log K \\
 -2895 &= -2.303 \times 8.314 \times 298 \times \log K
 \end{aligned}$$

$$\text{or } \log K = 0.5074$$

$$\text{or } K = \text{Antilog } (0.5074) = 3.22.$$

Question 5.

Write the Nernst Equation and calculate e.m.f. of the following cells at 298 K :

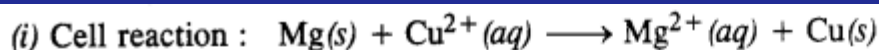
(i)  $\text{Mg(s)} \mid \text{Mg}^{2+} (0.001 \text{ M}) \parallel \text{Cu}^{2+} (0.0001 \text{ M}) \mid \text{Cu(s)}$  (C.B.S.E. Delhi 2008, 2013)

(ii)  $\text{Fe(s)} \mid \text{Fe}^{2+} (0.001 \text{ M}) \parallel \text{H}^+ (1 \text{ M}) \mid \text{H}_2(\text{g}) (1 \text{ bar}) \mid \text{Pt(s)}$

(iii)  $\text{Sn(s)} \mid \text{Sn}^{2+} (0.050 \text{ M}) \parallel \text{H}^+ (0.02 \text{ M}) \mid \text{H}_2(\text{g}) (1 \text{ bar}) \mid \text{Pt(s)}$  (C.B.S.E. Outside Delhi 2013, 2015)

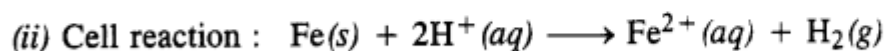
(iv)  $\text{Pt(s)} \mid \text{Br}_2(\text{l}) \mid \text{Br}^- (0.010 \text{ M}) \parallel \text{H}^+ (0.030 \text{ M}) \mid \text{H}_2(\text{g}) (1 \text{ bar}) \mid \text{Pt(s)}$

Answer:



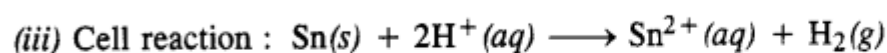
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]} = [0.34 - (-2.37)] - \frac{0.0591}{2} \log \frac{10^{-3}}{10^{-4}}$$

$$= 2.71 - 0.02955 = \mathbf{2.68 \text{ V}}$$



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{[\text{Fe}^{2+}]}{[\text{H}^{+}]^2} = [0 - (-0.44)] - \frac{0.0591}{2} \log \frac{10^{-3}}{(1)^2}$$

$$= 0.44 - \frac{0.0591}{2} \times (-3) = 0 = 0.44 + 0.0887 = \mathbf{0.523 \text{ V}}$$



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[\text{Sn}^{2+}]}{[\text{H}^{+}]^2} = [0 - (-0.14)] - \frac{0.0591}{2} \log \frac{0.05}{(0.02)^2}$$

$$= 0.14 - \frac{0.0591}{2} \times (2.097) = 0.14 - 0.0620 = \mathbf{0.08 \text{ V}}$$



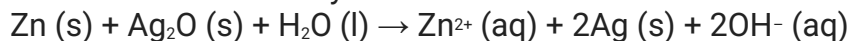
$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{2} \log \frac{1}{[\text{Br}^{-}]^2 [\text{H}^{+}]^2}$$

$$= (0 - 1.08) - \frac{0.0591}{2} \log \frac{1}{(0.01)^2 \times (0.03)^2} = -1.08 - \frac{0.0591}{2} \log(1.111 \times 10^7)$$

$$= -1.08 - \frac{0.0591}{2} (7.0457) = -1.08 - 0.208 = \mathbf{-1.288 \text{ V}}$$

### Question 6.

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



Determine  $E^{\circ}$  and  $\Delta G^{\circ}$  for the reaction. (C.B.S.B. Delhi 2005, Outside Delhi 2006 Supp., 2008, C.B.S.E. Sample Paper 2010)

Answer:

Zn is oxidised and  $\text{Ag}_2\text{O}$  is reduced (as

$\text{Ag}^{+}$  ions change into Ag)

$$E_{\text{cell}}^{\circ} = E_{\text{Ag}_2\text{O}/\text{Ag}}^{\circ} - E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = 0.344 - (-0.76)$$

$$= \mathbf{1.104 \text{ V}}$$

$$\Delta_r G^{\circ} = -nFE_{\text{cell}}^{\circ} = -2 \times 96500 \times 1.104$$

$$= \mathbf{-2.13 \times 10^5 \text{ J}}$$

### Question 7.

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



with concentration.

Answer:

The conductivity of a solution is defined as the conductance of a solution 1 cm in length and the area of cross-section  $\text{cm}^2$ . It is represented by  $\kappa$ .

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 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.

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

Question 8.

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

(C.B.S.E. Delhi 2008, 2013)

Answer:

$$\text{Conductivity } (\kappa) = 0.0248 \text{ S cm}^{-1} = 0.0248 \text{ ohm}^{-1} \text{ cm}^{-1}$$

$$\text{Molar concentration } (C) = 0.20 \text{ M} = 0.2 \text{ mol L}^{-1} = \frac{(0.2 \text{ mol})}{(1000 \text{ cm}^3)} = 2.0 \times 10^{-4} \text{ mol cm}^{-3}$$

$$\begin{aligned} \text{Molar conductivity } (\Lambda_m) &= \frac{\kappa}{C} = \frac{(0.0248 \text{ ohm}^{-1} \text{ cm}^{-1})}{(2.0 \times 10^{-4} \text{ mol cm}^{-3})} = 124 \text{ ohm}^{-1} \text{ mol}^{-1} \text{ cm}^2 \\ &= \mathbf{124 \text{ S mol}^{-1} \text{ cm}^2} \end{aligned}$$

Question 9.

The resistance of a conductivity cell containing 0.001 M KCl solution at 298 K is  $1500 \Omega$ . What is the cell constant if the conductivity of 0.001 M KCl solution at 298 K is  $0.146 \times 10^{-3} \text{ S cm}^{-1}$ ? (C.B.S.E. Outside Delhi 2007, 2008, 2013)

Answer:

$$\text{Resistance } (R) = 1500 \Omega$$

$$\text{Conductance } (C) = \frac{1}{R} = \frac{1}{1500 \text{ ohm}} = \frac{1}{1500} \text{ S}$$

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

$$\text{Cell constant } (l/a) = \frac{\text{Conductivity}}{\text{Conductance}} = \frac{(0.146 \times 10^{-3} \text{ S cm}^{-1})}{(1 / 1500 \text{ S})} = \mathbf{0.219 \text{ cm}^{-1}}$$



### Question 10.

The conductivity of sodium chloride solution at 298 K has been determined at different concentrations and results are given below :

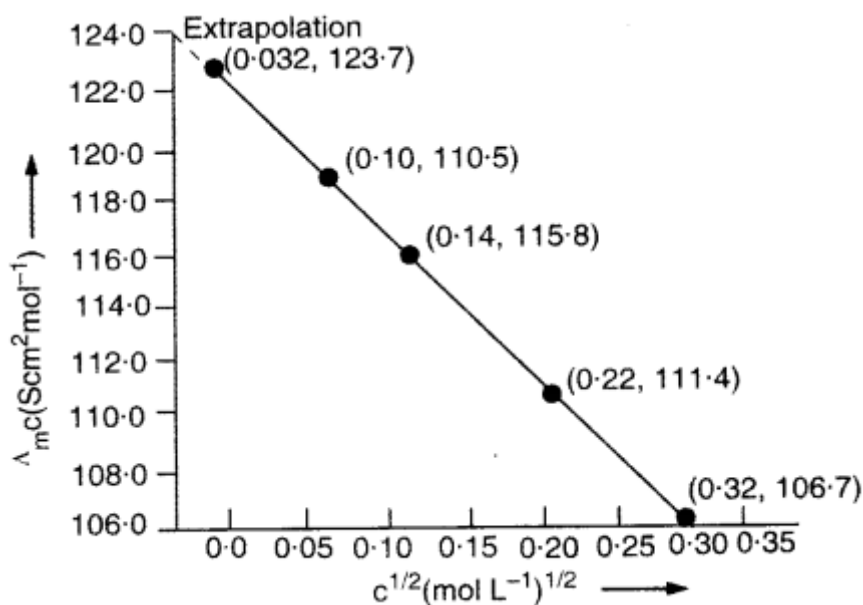
|                                                          |              |              |              |              |               |
|----------------------------------------------------------|--------------|--------------|--------------|--------------|---------------|
| <b>Concentration (mol L<sup>-1</sup>) :</b>              | <b>0.001</b> | <b>0.010</b> | <b>0.020</b> | <b>0.050</b> | <b>0.100</b>  |
| <b>Conductivity 10<sup>2</sup> × κ/Sm<sup>-1</sup> :</b> | <b>1.237</b> | <b>11.85</b> | <b>23.15</b> | <b>55.53</b> | <b>106.74</b> |

Calculate molar conductivity for all the concentrations and draw a plot between  $\Lambda_{cm}$  and  $c\sqrt{}$ . Find the value  $\Lambda_{\infty m}$  from the graph.

Answer:

$1\text{Scm}^{-1} = 100\text{Sm}^{-1} = 1$  (unit conversion factor)

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



$\Lambda_{\infty m}$  can be obtained on extrapolation to zero concentration along Y-axis. It is  $124.0\text{Scm}^2\text{mol}^{-1}$ .

### Question 11.

The conductivity of 0.00241 M acetic acid is  $7.896 \times 10^{-5} \text{ S cm}^{-1}$ . Calculate the molar conductivity. If  $\Lambda^{\circ}$  for acetic acid is  $390.5 \text{ S cm}^2 \text{ mol}^{-1}$ , what is its dissociation constant? (C.B.S.E. Delhi 2008, C.B.S.E. Outside Delhi 2016)

Answer:

**Step I.** Calculation of molar conductance ( $\Lambda_m^c$ )

$$\kappa = 7.896 \times 10^{-5} \text{ S cm}^{-1}; C = 0.00241 \text{ mol L}^{-1} = \frac{0.00241 \text{ mol}}{10^3 \text{ cm}^3} = 241 \times 10^{-8} \text{ mol cm}^{-3}$$

$$\Lambda_m^c = \frac{\kappa}{C} = \frac{(7.896 \times 10^{-5} \text{ S cm}^{-1})}{(241 \times 10^{-8} \text{ mol cm}^{-3})} = 32.76 \text{ S cm}^2 \text{ mol}^{-1}$$

**Step II.** Calculation of degree of dissociation of acetic acid.

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^0} = \frac{(32.76 \text{ S cm}^2 \text{ mol}^{-1})}{(390.5 \text{ S cm}^2 \text{ mol}^{-1})} = 0.084 = 8.4 \times 10^{-2}$$

**Step III.** Calculation of dissociation constant  $K_c$

$$K_c = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = \frac{C\alpha \times C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha}$$

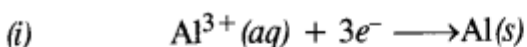
$$= \frac{(0.00241 \text{ mol L}^{-1}) \times (0.084)^2}{1 - 0.084} = 0.000185 \text{ mol L}^{-1} = 1.85 \times 10^{-5} \text{ mol L}^{-1}$$

Question 12.

How much charge is required for the reduction of :

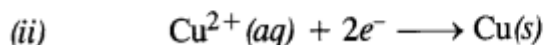
- (i) 1 mol of  $\text{Al}^{3+}$  to Al
- (ii) 1 mol of  $\text{Cu}^{2+}$  to Cu
- (iii) 1 mol of  $\text{MnO}_4^-$  to  $\text{Mn}^{2+}$ .

Answer:



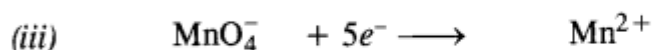
3 Faraday

$$\text{Charge} = 3 \times F = 3 \times 96500 \text{ C} = 2.895 \times 10^5 \text{ C}$$



2 Faraday

$$\text{Charge} = 2 \times F = 2 \times 96500 \text{ C} = 1.93 \times 10^5 \text{ C}$$



(O.N. of Mn = +7)

(O.N. of Mn = +2)

$$\text{Charge} = 5 \times F = 5 \times 96500 \text{ C} = 4.825 \times 10^5 \text{ C.}$$

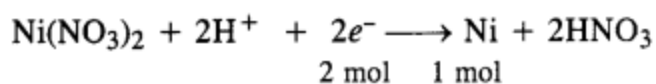
Question 13.

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:



The charge,  $Q$  on  $n$  moles of electrons is given by  $Q = nF$

Thus, charge required to deposit 1 mole of nickel,

$$Q = 2 \text{ mol} \times 96500 \text{ C mol}^{-1} = 193000 \text{ C} = 1.93 \times 10^5 \text{ C}$$

Quantity of electricity used = Current in amperes  $\times$  Time in seconds

$$= 5\text{A} \times (20 \times 60)\text{s} = 5 \times 20 \times 60 \text{ A s} = 6000 \text{ C} (\because \text{A s} = \text{C})$$

Molar mass of nickel =  $58.7 \text{ g mol}^{-1}$

$1.93 \times 10^5 \text{ C}$  of charge produce nickel = 1 mol =  $1 \text{ mol} \times 58.7 \text{ g mol}^{-1} = 58.7 \text{ g}$

$$\therefore 6000 \text{ C of charge produce nickel} = \frac{(58.7 \text{ g})}{(1.93 \times 10^5 \text{ C})} \times (6000 \text{ C}) = \mathbf{1.825 \text{ g.}}$$

Question 16.

Three electrolytic cells A, B, and C containing electrolytes zinc sulphate, silver nitrate, and copper sulphate respectively, were connected in series. A steady current of 1.50 ampere was passed through them until 1.45 g of silver was deposited at the cathode of cell B. How long did the current flow? What weight of copper and of zinc were deposited? (Atomic mass of Cu = 63.5; Zn = 65.3; Ag = 108) (C.B.S.E. Outside Delhi 2008, Jharkhand Board 2010)

Answer:



The charge  $Q$  on  $n$  moles of electrons is given by,  $Q = nF$

**Step I.** Calculation of time for the flow of current.

For reaction (ii),  $n = 1 \text{ mol}$

Thus,  $Q = 1 \text{ mol} \times 96500 \text{ C mol}^{-1} = 96500 \text{ C}$

Molar mass of silver =  $108 \text{ g mol}^{-1}$

Therefore, mass of 1 mole of silver =  $1 \text{ mol} \times 108 \text{ g mol}^{-1} = 108 \text{ g}$

108 g of Ag is deposited by electric charge = 96500 C

$\therefore$  1.45 g of Ag is deposited by electric charge =  $\frac{(96500 \text{ C})}{(108 \text{ g})} \times (1.45 \text{ g}) = 1295.6 \text{ C}$

Let 1.5 ampere of current be passed for time  $t$ , so that

quantity of electricity used =  $1.5 \text{ A} \times t$

Thus,  $1.5 \text{ A} \times t = 1295.6 \text{ C}$

$$t = \frac{(1295.6 \text{ C})}{(1.5 \text{ A})} = \frac{(1295.6 \text{ A s})}{(1.5 \text{ A})} = 863.7 \text{ s} \approx 14 \text{ min } 24 \text{ sec} \quad (\because \text{C} = \text{A s})$$

**Step II. Calculation of mass of zinc deposited.**

For reaction (i), No. of electron moles ( $n$ ) = 2 mol

Thus, Charge on 2 electron moles (Q) =  $2 \text{ mol} \times 96500 \text{ C mol}^{-1} = 193000 \text{ C} = 1.93 \times 10^5 \text{ C}$

Molar mass of Zn =  $65.3 \text{ g mol}^{-1}$

Therefore, mass of 1 mole of Zn =  $1 \text{ mol} \times 65 \text{ g mol}^{-1} = 65 \text{ g}$

$\therefore$   $1.93 \times 10^5 \text{ C}$  of charge deposit Zn = 65 g

$$1295.6 \text{ C of charge deposit zinc} = \frac{(65.3 \text{ g})}{(1.93 \times 10^5 \text{ C})} \times (1295.6 \text{ C}) = 0.438 \text{ g}$$

**Step III. Calculation of mass of copper deposited.**

For reaction (iii), No. of electron mols ( $n$ ) = 2 mol

Thus, Charge on 2 electron moles (Q) =  $2 \text{ mol} \times 96500 \text{ C mol}^{-1}$   
 $= 193000 \text{ C} = 1.93 \times 10^5 \text{ C}$

Molar mass of copper =  $63.5 \text{ g mol}^{-1}$

Therefore, mass of 1 mole of copper =  $1 \text{ mol} \times 63.5 \text{ g mol}^{-1} = 63.5 \text{ g}$

$\therefore$   $1.93 \times 10^5 \text{ C}$  of charge deposit copper = 63.5 g

$$1295.6 \text{ C of charge deposit copper} = \frac{(63.5 \text{ g})}{(1.93 \times 10^5 \text{ C})} \times (1295.6 \text{ C}) = 0.426 \text{ g}$$

Question 17.

Predict if the reaction between the following is feasible:

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

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

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

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

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

Given standard electrode potentials:

$$E^{\circ}_{\frac{1}{2} \text{I}_2, \text{I}^{-}} = +0.541 \text{ V},$$

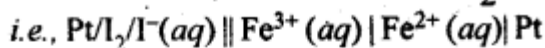
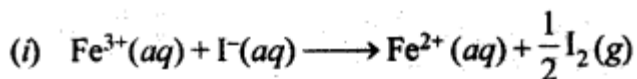
$$E^{\circ}_{\text{Cu}^{2+}, \text{Cu}} = +0.34 \text{ V},$$

$$E^{\circ}_{\frac{1}{2} \text{Br}_2, \text{Br}^{-}} = +1.09 \text{ V},$$

$$E^{\circ}_{\text{Ag}^{+}, \text{Ag}} = +0.80 \text{ V}, E^{\circ}_{\text{Fe}^{3+}, \text{Fe}^{2+}} = +0.77 \text{ V}.$$

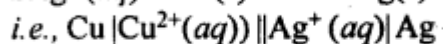
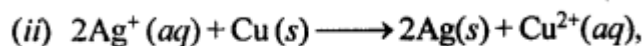
Answer:

The reaction is feasible if the EMF of the cell reaction is positive.



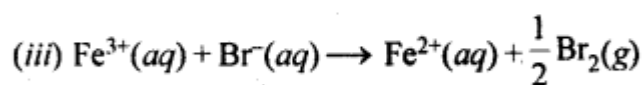
$$\therefore E_{\text{cell}}^{\circ} = E_{\text{Fe}^{3+}, \text{Fe}^{2+}}^{\circ} - E_{1/2\text{I}_2, \text{I}^{-}}^{\circ}$$

$$= 0.77 - 0.54 = 0.23 \text{ V (Feasible).}$$



$$E_{\text{cell}}^{\circ} = E_{\text{Ag}^{+}, \text{Ag}}^{\circ} - E_{\text{Cu}^{2+}, \text{Cu}}^{\circ}$$

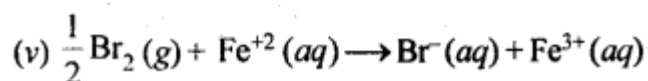
$$= 0.80 - 0.34 = 0.46 \text{ V (Feasible).}$$



$$E_{\text{cell}}^{\circ} = 0.77 - 1.09 = -0.32 \text{ V (Not feasible).}$$



$$E_{\text{cell}}^{\circ} = 0.77 - 0.80 = -0.03 \text{ V (Not feasible).}$$



$$E_{\text{cell}}^{\circ} = 1.09 - 0.77 = 0.32 \text{ V (Feasible).}$$

Question 18.

Predict the products of electrolysis of each of the following :

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

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

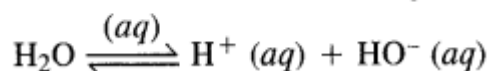
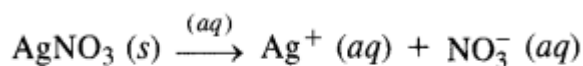
(iii) A dilute solution of  $\text{H}_2\text{SO}_4$  using platinum electrodes. (C.B.S.E. Outside Delhi 2007)

(iv) An aqueous solution of  $\text{CuCl}_2$  using platinum electrodes. (C. B. S. E. Sample Paper 2010)

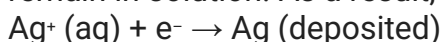
Answer:

**(i) An aqueous solution of  $\text{AgNO}_3$  using silver electrodes :**

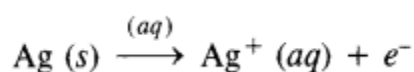
Both  $\text{AgNO}_3$  and water will ionise in aqueous solution



At cathode:  $\text{Ag}^{+}$  ions with less discharge potential are reduced in preference to  $\text{H}^{+}$  ions which will remain in solution. As a result, silver will be deposited at cathode.



At anode: An equivalent amount of silver will be oxidised to  $\text{Ag}^{+}$  ions by releasing electrons.



**(ii) An aqueous solution of  $\text{AgNO}_3$  using platinum electrodes:**

In this case, the platinum electrodes are the non-attackable electrodes. On passing current the following changes will occur at the electrodes.

**At cathode:**  $\text{Ag}^{+}$  ions will be reduced to Ag which will get deposited at the cathode.

**At anode:** Both  $\text{NO}_3^{-}$  and  $\text{OH}^{-}$  ions will migrate. But  $\text{OH}^{-}$  ions with less discharge potential will be



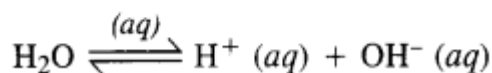
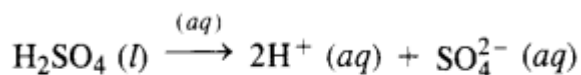
oxidised in preference to  $\text{NO}_3^-$  ions which will remain in solution.



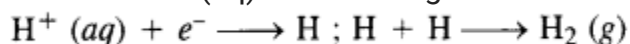
Thus, as a result of electrolysis, silver is deposited on the cathode while  $\text{O}_2$  is evolved at the anode. The solution will be acidic due to the presence of  $\text{HNO}_3$ .

**(iii) A dilute solution of  $\text{H}_2\text{SO}_4$  using platinum electrodes:**

On passing current, both acid and water will ionise as follows:

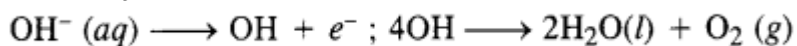


**At cathode:**  $\text{H}^+$  (aq) ions will migrate to the cathode and will be reduced to  $\text{H}_2$ .



Thus,  $\text{H}_2$  (g) will be evolved at the cathode.

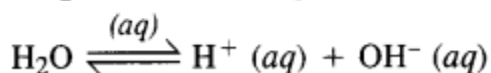
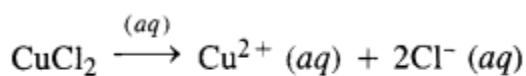
**At anode:**  $\text{OH}^-$  ions will be released in preference to  $\text{SO}_4^{2-}$  ions because their discharge potential is less. They will be oxidized as follows:



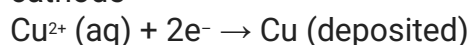
Thus,  $\text{O}_2$  (g) will be evolved at the anode. The solution will be acidic and will contain  $\text{H}_2\text{SO}_4$ .

**(iv) An aqueous solution of  $\text{CuCl}_2$  using platinum electrodes :**

The electrolysis proceeds in the same manner as discussed in the case of  $\text{AgNO}_3$  solution. Both  $\text{CuCl}_2$  and  $\text{H}_2\text{O}$  will ionise as follows :



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



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



Thus,  $\text{Cl}_2$  will be evolved at the anode.

We hope the NCERT Solutions for Class 12 Chemistry Chapter 3 Electrochemistry help you. If you have any query regarding NCERT Solutions for Class 12 Chemistry Chapter 3 Electrochemistry, drop a comment below and we will get back to you at the earliest.