CBSE Test Paper-03

Class 12 Chemistry (Electrochemistry)

1. Molar conductivity denoted by the symbol Λm is related to the conductivity of the solution by the equation (κ is the conductivity and c the concentration)
   a. $\Lambda m = \frac{c}{k}$
   b. $\Lambda m = \frac{k}{c}$
   c. $\Lambda m = kc$
   d. $\Lambda m = \frac{2k}{c}$
2. 96500 coulomb deposit 107.9 g of Ag from its solution. If $e = 1.6 \times 10^{-19}$coulomb, calculate the number of electrons per mole of electrons.
   a. 6.02×10^{23}
   b. 96500
   c. 1.6×10^{-19}
   d. 6.02×10^{-23}
3. Resistance of 0.2 M solution of an electrolyte is 50Ω. The specific conductance of the solution is 1.3s m^{-1}. If resistance of the 0.4 M solution of the same electrolyte is 260Ω, its molar conductivity is
   a. $6.25 \times 10^{-4} S m^2 mol^{-1}$
   b. $625 \times 10^{-4} S m^2 mol^{-1}$
   c. $6250 \ S \ m^2 mol^{-1}$
   d. $62.5 \ S \ m^2 mol^{-1}$
4. Standard electrode potential is measured taking the concentrations of all the species involved in a half-cell is
   a. 2M
   b. 3M
   c. 1 M
   d. 4 M
5. Three electrolytic cells A,B,C containing solutions of $ZnSO_4$, $AgNO_3$ and $CuSO_4$, respectively are connected in series. A steady current of 1.5 amperes was passed through them until 1.45 g of silver deposited at the cathode of cell. How long did the current flow? What mass of copper and zinc were deposited?

a. 823s, Copper 0.487g, Zinc 0.437 g
b. 863s, Copper 0.426g, Zinc 0.438 g
c. 763s, Copper 0.403g, Zinc 0.437 g
d. 800s, Copper 0.452g, Zinc 0.437g

6. Give some uses of electrochemical cells?

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

8. What is the reference electrode in determining the standard electrode potential?

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

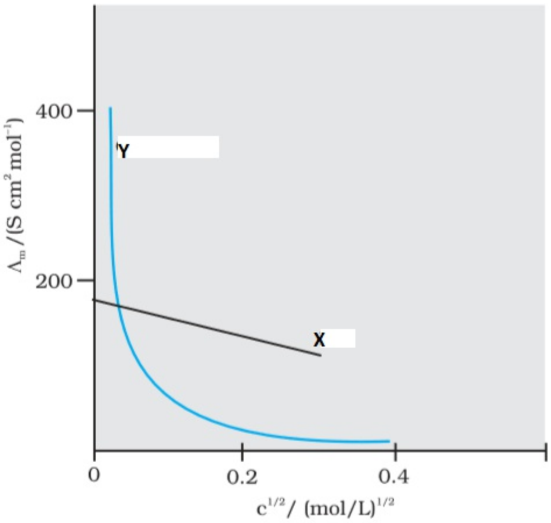
10. Conductivity of 0.001 M acetic acid is $4 \times 10^{-5} Scm^{-1}$. Calculate its molar conductivity? if its molar conductivity at infinite dilution is 390 S cm^2 mol^{-1}, what is its dissociation constant?

11. State Faraday's Laws of electrolysis?

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

13. a. Write the formulation for the galvanic cell in which the reaction Cu(s) + $2Ag^+$ (aq) $\rightarrow$ Cu^{2+}(aq) + 2Ag(s) takes place.
    Identify the cathode and the anode reactions in it.
    b. Write Nernst equation and calculate the emf of the following cell: Sn(s) | Sn^{2+} (0.04M) || H^+(0.02M) | H_2(g)|Pt(s)
    $\left(Given \;\; E^{\ominus}{}_{Sn^{2+}/Sn} = -0.14V\right)$
14. How much charge is required for the following reductions:
    i. 1 mol of Al^{3+} to Al.
    ii. 1 mol of Cu^{2+} to Cu.
    iii. 1 mol of MnO_4^- to Mn^{2+}
15. The following curve is obtained when molar conductivity, Λ_m is plotted against the square root of concentration, $C^{1/2}$ along y and x-axis respectively for the two

electrolytes X and Y.



i. What can you say about the nature of these two electrolytes?
ii. How do you account for the increase in Λ_m for the electrolytes X and Y with dilution?
iii. How can you determine Λ_m^∞ for these electrolytes?

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Class 12 Chemistry (Electrochemistry)
Solutions

1. b. $\Lambda m = \frac{k}{c}$

   Explanation: Molar conductivity is conducting power of all the ions produced by dissolving 1 mole of electrolyte in solution.

2. a. 6.02×10^{23}

   Explanation: 6.02×10^{23}

   1 mol of electrons contains avagadro number (6.02 $\times$ 10^{23}) of electrons

3. a. $6.25 \times 10^{-4} Sm^2 mol^{-1}$

   Explanation: κ = G $\times$cell constant and G= 1/R. For 0.2M solution, R = 50 ohms, k= $1.3Sm^{-1}$, cell constant l/A = 50 × 1.3 = $65m^{-1}$

   $\Lambda m = \frac{k}{c}$ for 0.4M solution , molar conductivity = ?, R = 260ohms, cell constant = $65m^{-1}$ so K = 65/260 = 0.25 Sm^{-1}

   molar conductivity = K × 1000/M = 0.25/0.4 × 1000 = 6.25 × 10^{-4} Sm^2/mol

4. c. 1 M

   Explanation: Under standard condition that is at 298K, the molar concentration of every species taken in each half cell is 1 M and if gases are involved the partial pressure of the gas is taken as 1 atm.

5. b. 863s, Copper 0.426g, Zinc 0.438 g

   Explanation: Ag^+ + e = Ag 96500C of current deposits 108g of silver . so 1.45g of silver would be deposited by = 96500 $\times$ 1.45/108 = 1295.6C

   Q = I $\times$ t 1295.6C = 1.5 $\times$ t therefore t = 863s
   $W = \frac{(E \times I \times t)}{96500}$ Equivalent mass of copper = 31.75, equivalent mass of zinc = 32.65
   mass of zinc deposited = 32.65 $\times$ 1.5× 863/96500 = 0.438g
   mass of copper deposited = 31.75 $\times$ 1.5× 863/96500 = 0.426g

6. An electrochemical cell is a device which is capable of producing electrical energy from chemical reactions or making chemical reactions easy through the introduction of electrical energy. Electrochemical cells are used for determining the:
   a. pH of solutions
   b. solubility product and equilibrium constant
   c. in potentiometric titrations

7. A paste of NH_4Cl, MnO_2 and Carbon.

8. Standard hydrogen electrode (SHE) is used as reference electrode in determining the standard electrode potential because its reduction and oxidation potential value is 0.0 volt.

9. Given that,

   $\kappa = 0.0248\ S\ cm^{-1}$ and $c = 0.20\ M$

   Formula of molar conductivity,

   $\Lambda_m = (k \times 1000)/M$

   $\Lambda_m = (0.0248 \times 1000)/0.2$

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

10. $\wedge_m^c = \frac{k \times 1000}{C}$

    $= \frac{4 \times 10^{-5} \times 1000}{0.001} S\ cm^2/mol = 40\ Scm^2/mol$

    Degree of dissociation,

    $\alpha = \frac{\wedge^C}{\wedge^\infty} = \frac{40}{390} = 0.10256$

    For the dissociation of acetic acid, CH_3COOH

    $$\begin{array}{llll} & CH_3COOH \rightleftharpoons & CH_3COO^- + & H^+ \\ Initial\ (conc.) & C & 0 & 0 \\ After\ time,\ t & (C-C\alpha) & C\alpha & C\alpha \end{array}$$

    Dissociation constant,

    $K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$

    $= \frac{C\alpha \cdot C\alpha}{C-C\alpha} = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$

    $K_a = \frac{0.001 \times (0.103)^2}{(1-0.103)}$

    $= \frac{1.061 \times 10^{-5}}{0.897} = 1.18 \times 10^{-5}$

11. **Faraday's Laws of electrolysis :**

    First Law: The amount of chemical reaction which occurs at any electrode during electrolysis by a current is proportional to the quantity of electricity passed through the electrolyte.

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

12. Conductivity of a solution is defined as the conductance of a solution of 1 cm length and having 1 sq. cm as the area of cross section.

    Specific conductivity

    $\kappa = \frac{1}{\rho} = \frac{1}{ohm\,cm}$

    = $ohm^{-1}cm^{-1}$

    In terms of SI units = sm^{-1}

    Molar conductivity: It is defined as the conducting power of all the ions produced by one gram mole of an electrolyte in a solution and it is denoted by Λ_m.

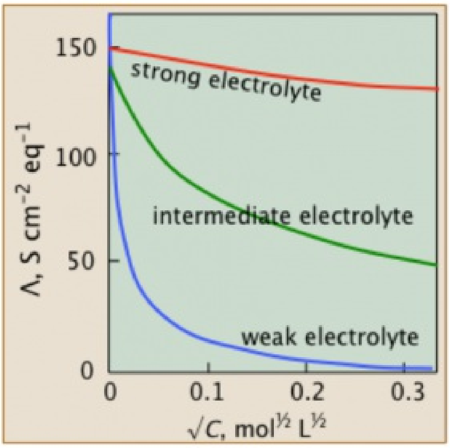
    $\Lambda_m = \frac{k}{C} \times 1000 S\,cm^2 mol^{-1}$

    Where 'k' is electrolytic conductivity of solution and 'C' is concentration of the solution expressed in mol L^{-1}

    The unit of Λ_m = $1Sm^2mol^{-1}$

    Variation of conductivity and molar conductivity with concentration: Both conductivity and molar conductivity change with change in concentration of electrolyte. Conductivity always decreases with decrease in concentration both for weak as much as strong electrolytes. It is because number of ions per unit volume that carry the current in a solution decreases on dilutions. Molar conductivity increases with decreases in concentration. This is because both number of ions as well as mobility of ions increases with dilution. When concentration approaches zero, the molar conductivity is known as limiting molar conductivity or molar conductivity at infinite dilution.

    The variation in Λ_m with concentration for strong and weak electrolytes.



13. a. We have

$E^{\Theta}_{(Cu^{2+}/Cu)} = 0.34V$ and $E^{\Theta}_{(Ag^{+}/Ag)} = 0.80V$

Standard emf of Cu is less than Ag, therefore it is strong reducing agent and is oxidised. Therefore Cu acts as Anode and Ag acts as Cathode.

Half cell reactions are:

At Cathode (Reduction):

$2Ag^{+}(aq) + 2e^{-} \rightarrow 2Ag(s)$

At Anode (Oxidation):

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

b. The reactions are :

At Anode:

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^{-}$

At Cathode:

$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$

Full cell reaction:

$Sn(s) + 2H^{+}(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$

Standard emf of the cell is:

$E^0_{cell} = E^0_{H^+/H_2} - E^0_{Sn^{2+}/Sn}$

= 0 - (-0.14)V

= + 0.14V

For this reaction n=2 moles of electrons. Using Nernst equation,

$E_{cell} = 0.14 - \frac{0.0591}{2} \log \frac{[Sn^2+]}{[H^+]^2}$

$= 0.14 - \frac{0.0591}{2} \log \frac{0.04}{(0.02)^2}$

$= 0.14 - \frac{0.0591}{2} \log \frac{4}{100} \times \frac{100}{2} \times \frac{100}{2}$

= 0.14 V - 0.0591 V

= 0.0809 V

14. i. $Al^{3+} + 3e^- \rightarrow Al$

Therefore, Required charge = 3 F

$= 3 \times 96487C$

= 289461 C

ii. $Cu^{2+} + 2e^- \rightarrow Cu$

Therefore, Required charge = 2 F

$= 2 \times 96487C$

= 192974 C

iii. $MnO_4^- \rightarrow Mn^{2+}$

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

Therefore, Required charge = 5 F

$= 5 \times 96487C$

= 482435 C

15. i. Electrolyte X is a strong electrolyte and Y is a weak electrolyte.

ii. Molar conductivity,Λ_m of X (strong electrolysis) increases slowly with dilution. This is because interionic forces of attraction decreases on dilution, although the number of ions remain the same. As a result ions move freely and hence Λ_m increases with dilution. On the other hand, for Y (weak electrolyte) Λ_m increases sharply with dilution.This is because degree of dissociation increases on dilution resulting in greater number of ions on dilution. Hence Λ_m increases.

iii. For X, Λ_m^∞ can be obtained by extrapolation to zero concentration.