

Solved Examples On Electrochemistry

Example 1. Find the charge in coulomb on 1 g-ion of

Solution: Charge on one ion of N^{3-}
 $= 3 \times 1.6 \times 10^{-19}$ coulomb

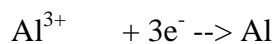
Thus, charge on one g-ion of N^{3-}

$$= 3 \times 1.6 \times 10^{-19} \times 6.02 \times 10^{23}$$

$$= 2.89 \times 10^5 \text{ coulomb}$$

Example 2. How much charge is required to reduce (a) 1 mole of Al^{3+} to Al and (b) 1 mole of Mn^{2+} to Mn^{2+} ?

Solution: (a) The reduction reaction is



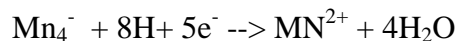
1 mole 3 mole

Thus, 3 mole of electrons are needed to reduce 1 mole of Al^{3+} .

$$Q = 3 \times F$$

$$= 3 \times 96500 = 289500 \text{ coulomb}$$

(b) The reduction is



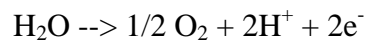
1 mole 5 mole

$$Q = 5 \times F$$

$$= 5 \times 96500 = 482500 \text{ coulomb}$$

Example 3. How much electric charge is required to oxidise (a) 1 mole of H_2O to O_2 and (b) 1 mole of FeO to Fe_2O_3 ?

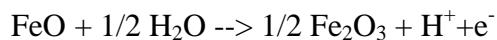
Solution: (a) The oxidation reaction is



1 mole 2 mole

$$\begin{aligned} Q &= 2 \times F \\ &= 2 \times 96500 = 193000 \text{ coulomb} \end{aligned}$$

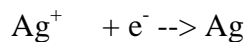
(b) The oxidation reaction is



$$Q = F = 96500 \text{ coulomb}$$

Example 4. Exactly 0.4 faraday electric charge is passed through three electrolytic cells in series, first containing AgNO_3 , second CuSO_4 and third FeCl_3 solution. How many gram of each metal will be deposited assuming only cathodic reaction in each cell?

Solution: The cathodic reactions in the cells are respectively.



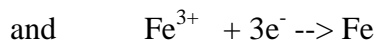
1 mole 1 mole

108 g 1 F



1 mole 2 mole

63.5 g 2 F



1 mole 3 mole

$$56 \text{ g} \quad 3 \text{ F}$$

Hence, $\text{Ag deposited} = 108 \times 0.4 = 43.2 \text{ g}$

$$\text{Cu deposited} = 63.5/2 \times 0.4 = 12.7 \text{ g}$$

and $\text{Fe deposited} = 56/3 \times 0.4 = 7.47 \text{ g}$

Example 5. An electric current of 100 ampere is passed through a molten liquid of sodium chloride for 5 hours. Calculate the volume of chlorine gas liberated at the electrode at NTP.

Solution: The reaction taking place at anode is



$$71.0 \text{ g} \quad 71.0 \text{ g} \quad 2 \times 96500 \text{ coulomb}$$

$$1 \text{ mole}$$

$$Q = I \times t = 100 \times 5 \times 60 \text{ coulomb}$$

The amount of chlorine liberated by passing $100 \times 5 \times 60 \times 60$ coulomb of electric charge.

$$= 1/(2 \times 96500) \times 100 \times 5 \times 60 \times 60 = 9.3264 \text{ mole}$$

$$\text{Volume of Cl}_2 \text{ liberated at NTP} = 9.3264 \times 22.4 = 208.91 \text{ L}$$

Example 6. A 100 watt, 100 volt incandescent lamp is connected in series with an electrolytic cell containing cadmium sulphate solution. What mass of cadmium will be deposited by the current flowing for 10 hours?

Solution: We know that

$$\text{Watt} = \text{ampere} \times \text{volt}$$

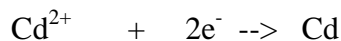
$$100 = \text{ampere} \times 110$$

$$\text{Ampere} = 100/110$$

$$\text{Quantity of charge} = \text{ampere} \times \text{second}$$

$$= 100/110 \times 10 \times 60 \times 60 \text{ coulomb}$$

The cathodic reaction is



$$112.4 \text{ g} \quad 2 \times 96500 \text{ C}$$

Mass of cadmium deposited by passing $100/110 \times 10 \times 60 \times 60$

$$\text{Coulomb charge} = 112.4 / (2 \times 96500) \times 100 / 110 \times 10 \times 60 \times 60 = 19.0598 \text{ g}$$

Example 7. In an electrolysis experiment, a current was passed for 5 hours through two cells connected in series. The first cell contains a solution gold salt and the second cell contains copper sulphate solution. 9.85 g of gold was deposited in the first cell. If the oxidation number of gold is +3, find the amount of copper deposited on the cathode in the second cell. Also calculate the magnitude of the current in ampere.

Solution: We know that

$$(\text{Mass of Au deposited}) / (\text{Mass of Cu deposited}) = (\text{Eq. mass of Au}) / (\text{Eq. Mass of Cu})$$

$$\text{Eq. mass of Au} = 197/3; \text{Eq. mass of Cu} = 63.5/2$$

Mass of copper deposited

$$= 9.85 \times 63.5/2 \times 3/197 \text{ g} = 4.7625 \text{ g}$$

Let Z be the electrochemical equivalent of Cu.

$$E = Z \times 96500$$

$$\text{or } Z = E/96500 = 63.5 / (2 \times 96500)$$

Applying $W = Z \times I \times t$

$$T = 5 \text{ hour} = 5 \times 3600 \text{ second}$$

$$4.7625 = 63.5 / (2 \times 96500) \times I \times 5 \times 3600$$

$$\text{or } I = (4.7625 \times 2 \times 96500) / (63.5 \times 5 \times 3600) = 0.0804 \text{ ampere}$$

Example 8. How long has a current of 3 ampere to be applied through a solution of silver nitrate to coat a metal surface of 80 cm² with 0.005 cm thick layer? Density of silver is 10.5 g/cm³.

Solution: Mass of silver to be deposited

$$= \text{Volume} \times \text{density}$$

$$= \text{Area} \times \text{thickness} \times \text{density}$$

Given: Area = 80 cm², thickness = 0.0005 cm and density = 10.5 g/cm³

$$\text{Mass of silver to be deposited} = 80 \times 0.0005 \times 10.5$$

$$= 0.42 \text{ g}$$

Applying to silver $E = Z \times 96500$

$$Z = 108/96500 \text{ g}$$

Let the current be passed for r seconds.

We know that

$$W = Z \times I \times t$$

$$\text{So, } 0.42 = 108/96500 \times 3 \times t$$

$$\text{or } t = (0.42 \times 96500)/(108 \times 3) = 125.09 \text{ second}$$

Example 9. What current strength in ampere will be required to liberate 10 g of chlorine from sodium chloride solution in one hour?

Solution: Applying $E = Z \times 96500$ (E for chlorine = 35.5)

$$35.5 = Z \times 96500$$

$$\text{or } Z = 35.5/96500 \text{ g}$$

Now, applying the formula

$$W = Z \times I \times t$$

Where $W = 10 \text{ g}$, $Z = 35.5/96500$ $t = 60 \times 60 = 3600$ second

$$I = 10 \times 96500 / 35.5 \times 96500 = 7.55 \text{ ampere}$$

Example 10. 0.2964 g of copper was deposited on passage of a current of 0.5 ampere for 30 minutes through a solution of copper sulphate. Calculate the atomic mass of copper. (1 faraday = 96500 coulomb)

Solution: Quantity of charge passed

$$0.5 \times 30 \times 60 = 900 \text{ coulomb}$$

$$900 \text{ coulomb deposit copper} = 0.2964 \text{ g}$$

$$96500 \text{ coulomb deposit copper} = 0.2964/900 \times 96500 = 31.78 \text{ g}$$

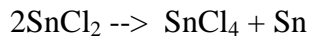
Thus, 31.78 is the equivalent mass of copper.

$$\text{At. mass} = \text{Eq. mass} \times \text{Valency}$$

$$= 31.78 \times 2 = 63.56$$

Example 11. 19 g of molten SnCl_2 is electrolysed for some time using inert electrodes until 0.119 g of Sn is deposited at the cathode. No substance is lost during electrolysis. Find the ratio of the masses of SnCl_2 : SnCl_4 after electrolysis.

Solution: The chemical reaction occurring during electrolysis is



$$2 \times 190 \text{ g} \quad 261 \text{ g} \quad 119 \text{ g}$$

119 g of Sn is deposited by the decomposition of 380 g of SnCl_2

So, 0.119 g of Sn of SnCl_2 is deposited by the decomposition of

$$380/119 \times 0.119 = 0.380 \text{ g of SnCl}_2$$

Remaining amount of $\text{SnCl}_2 = (19 - 0.380) = 18.62 \text{ g}$

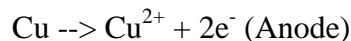
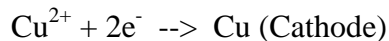
380 g of SnCl_2 produce = 261 g of SnCl_4

So 0.380 g of SnCl_2 produce = $261/380 \times 0.380 = 0.261$ g of SnCl_4

Thus, the ratio $\text{SnCl}_2 : \text{SnCl}_4 = 18.2/0.261$, i.e., 71.34 : 1

Example 12. A current of 2.68 ampere is passed for one hour through an aqueous solution of copper sulphate using copper electrodes. Calculate the change in mass of cathode and that of the anode. (At. mass of copper = 63.5).

Solution: The electrode reactions are:



Thus, cathode increases in mass as copper is deposited on it and the anode decreases in mass as copper from it dissolves.

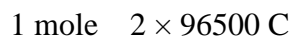
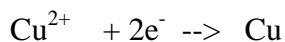
Charge passed through cell = $2.68 \times 60 \times 60$ coulomb

Copper deposited or dissolved = $63.5 / (2 \times 96500) \times 2.68 \times 60 \times 60$
= 3.174 g

Increase in mass of cathode = Decrease in mass of anode = 3.174 g

Example 13. An ammeter and a copper voltameter are connected in series through which a constant current flows. The ammeter shows 0.52 ampere. If 0.635 g of copper is deposited in one hour, what is the percentage error of the ammeter? (At. mass of copper = 63.5)

Solution : The electrode reaction is:



63.5 g of copper deposited by passing charge = 2×96500 Coulomb

0.635 g of copper deposited by passing charge

$$=(2 \times 96500) / 63.5 \times 0.653 \text{ coulomb}$$

$$= 2 \times 965 \text{ coulomb}$$

$$= 1930 \text{ coulomb}$$

We know that

$$Q = I \times t$$

$$1930 = I \times 60 \times 60$$

$$I = 1930 / 3600 = 0.536 \text{ ampere}$$

$$\text{Percentage error} = ((0.536 - 0.52) / 0.536) \times 100 = 2.985$$

Example 14. A current of 3.7 ampere is passed for 6 hours between platinum electrodes in 0.5 litre of a 2 M solution of $\text{Ni}(\text{NO}_3)_2$. What will be the molarity of the solution at the end of electrolysis? What will be the molarity of solution if nickel electrodes are used? (1 F = 96500 coulomb; Ni = 58.7)

Solution: The electrode reaction is



1 mole $2 \times 96500 \text{ C}$

Quantity of electric charge passed

$$= 3.7 \times 6 \times 60 \times 60 \text{ coulomb} = 79920 \text{ coulomb}$$

Number of moles of $\text{Ni}(\text{NO}_3)_2$ decomposed or nickel deposited = $(1.0 - 0.4140) = 0.586$

Since 0.586 moles are present in 0.5 litre,

Molarity of the solution = $2 \times 0.586 = 1.72 \text{ M}$

When nickel electrodes are used, anodic nickel will dissolve and get deposited at the cathode. The molarity of the solution will, thus, remain unaffected.

Example 15: An acidic solution of Cu^{2+} salt containing 0.4 g of Cu^{2+} is electrolysed until all the copper is deposited. The electrolysis is continued for seven more minutes with volume of solution kept at 100 mL and the current at 1.2 amp. Calculate the gases evolved at NTP during the entire electrolysis.

Solution: 0.4 g of $\text{Cu}^{2+} = 0.4/31.75 = 0.0126$ g equivalent

At the same time, the oxygen deposited at anode

$$= 0.0126 \text{ g equivalent}$$

$$= 8/32 \times 0.0126 = 0.00315 \text{ g mol}$$

After the complete deposited of copper, the electrolysis will discharge hydrogen at cathode and oxygen at anode.

The amount of charge passed = $1.2 \times 7 \times 60 = 504$ coulomb

So, Oxygen liberated = $1/96500 \times 504 = 0.00523$ g equivalent

$$= 8/32 \times 0.00523 = 0.001307 \text{ g mole}$$

Hydrogen liberated = 0.00523 g equivalent

$$= 1/2 \times 0.00523 = 0.00261 \text{ g mole}$$

Total gases evolved = $(0.00315 + 0.001307 + 0.00261)$ g mole

$$= 0.007067 \text{ g mole}$$

Volume of gases evolved at NTP

$$= 22400 \times 0.007067 \text{ mL}$$

$$= 158.3 \text{ mL}$$

Example 16: An acidic solution of Cu^{2+} salt containing 0.4 g of Cu^{2+} is electrolysed until all the copper is deposited. The electrolysis is continued for seven more minutes with volume of solution kept at 100 mL and the current at 1.2 amp. Calculate the gases evolved at NTP during the entire electrolysis.

Solution: Amount of charge passed = 1.70×230 coulomb

$$\begin{aligned} \text{Amount of actual charge passed} &= 90/100 \times 1.70 \times 230 \\ &= 351.9 \text{ coulomb} \end{aligned}$$

No. of moles of Zn deposited by passing 351.9 coulomb of charge

$$= 1/(2 \times 96500) \times 351.9 = 0.000182$$

Molarity of Zn^{2+} ions after deposition of zinc

$$\begin{aligned} &= [0.160 - (0.000182 \times 1000) / 300] \text{M} \\ &= 0.154 \text{ M} \end{aligned}$$

Example 18: Calculate the electricity that would be required to reduce 12.3 g of nitrobenzene to aniline, if the current efficiency for the process is 50 per cent. If the potential drop across the cell is 3.0 volt, how much energy will be consumed?

Solution: The reduction reaction is



$$123 \text{ g} \quad 6\text{g}$$

$$1 \text{ mole} \quad 3 \text{ mole}$$

Hydrogen required for reduction of $12.3/123$ or 0.1 mole of nitrobenzene = $0.1 \times 3 = 0.3$ mole

Amount of charge required for liberation of 0.3 mole of hydrogen = $2 \times 96500 \times 0.3 = 57900$ coulomb

Actual amount of charge required as efficiency is 50%

$$= 2 \times 57900 = 115800 \text{ coulomb}$$

Energy consumed = $115800 \times 3.0 = 347400$ J

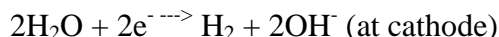
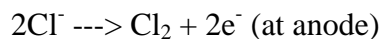
$$= 347.4 \text{ kJ}$$

Example 18: An aqueous solution of sodium chloride on electrolysis gives $\text{H}_2(\text{g})$, $\text{Cl}_2(\text{g})$ and NaOH according to the reaction:



A direct current of 25 ampere with a current efficiency 62% is passed through 20 L of NaCl solution (20% by mass). Write down the reactions taking place at the anode and cathode. How long will it take to produce 1 kg of Cl_2 ? What will be the molarity of the solution with respect to hydroxide ion? Assume no loss due to evaporation.

Solution: Reactions at anode and cathode are



$$1 \text{ kg of } \text{Cl}_2 = 1000/71.0 = 14.08 \text{ mole}$$

Charge to produce one mole of $\text{Cl}_2 = 2 \times 96500$ coulomb

Charge to produce 14.08 mole of $\text{Cl}_2 = 2 \times 96500 \times 14.08$ coulomb

Effective current = $62/100 \times 25.0 = 15.5$ ampere

Time = charge/current = $(2 \times 96500 \times 14.08)/15.5$

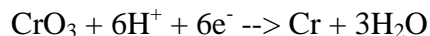
$$= 175318.7 \text{ second} = 48.699 \text{ hour}$$

OH^- ions produced = $2 \times$ mole of Cl_2

$$= 2 \times 14.08 = 28.16 \text{ mole}$$

Molarity Mole/Volume = $28.16/20 = 1.408 \text{ M}$

Example 19: Chromium metal can be plated out from an acidic solution containing CrO_3 according to the following reaction:



Solution: $\text{CrO}_3 + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{Cr} + 3\text{H}_2\text{O}$

$$6 \times 96500 \text{ C} \quad 1 \text{ mole}$$

$$52 \text{ g.}$$

Mass of chromium plated out by 24000 coulomb charge

$$= \frac{52}{6} \times 96500 \times \frac{24000}{96500} = 2.155 \text{ g}$$

Charge required for plating out 1.5 g of chromium

$$= \frac{6 \times 96500}{52} \times 1.5 = 16701.92 \text{ coulomb}$$

Time = charge/current = 1336.15 second

$$= 22.27 \text{ minute}$$

Example 20: After electrolysis of a sodium chloride solution with inert electrodes for a certain period of time, 600 mL of the solution was left which was found to be 1 N in NaOH. During the same period 31.75 g of copper was deposited in the copper voltameter in series with the electrolytic cell. Calculate the percentage theoretical yield of NaOH obtained.

Solution: Equivalent mass of NaOH = $\frac{40}{1000} \times 600 = 24 \text{ g}$

$$\text{Amount of NaOH formed} = \frac{40}{1000} \times 600 = 24 \text{ g}$$

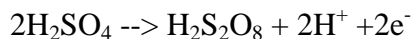
$$31.75 \text{ g of Cu} = 1 \text{ g equivalent of Cu.}$$

During the same period, 1 g equivalent of NaOH should have been formed.

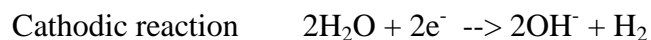
$$1 \text{ g equivalent of NaOH} = 40 \text{ g}$$

$$\% \text{ yield} = \frac{24}{40} \times 100 = 60$$

Example 21: Peroxy disulphuric acid ($\text{H}_2\text{S}_2\text{O}_8$) can be prepared by electrolytic oxidation of H_2SO_4 as:



Oxygen and hydrogen are by products. In such an electrolysis 9.72 litre of H_2 and 2.35 litre of O_2 were generated at NTP. What is the mass of peroxy disulphuric acid formed?

Solution:

Total equivalent of $\text{H}_2\text{S}_2\text{O}_8$ + equivalent of oxygen = Equivalent of H_2

$$9.72 \text{ litre } \text{H}_2 = 9.72/11.2 = 0.868 \text{ equivalent}$$

$$2.35 \text{ litre } \text{O}_2 = 2.35/5.6 = 0.42 \text{ equivalent}$$

$$\text{Equivalent of } \text{H}_2\text{S}_2\text{O}_8 = (0.868 - 0.420)$$

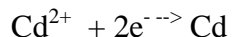
$$= 0.448$$

$$\text{Mass of } \text{H}_2\text{S}_2\text{O}_8 = 0.448 \times 194/2 = 43.456 \text{ g.}$$

Example 22: Cadmium amalgam is prepared by electrolysis of a solution of CdCl_2 using a mercury cathode. Find how long a current of 5 ampere should be passed in order to prepare 12% Cd-Hg amalgam on a cathode of 2 g mercury. At mass of Cd = 112.40.

Solution: 2 g Hg require Cd to prepare 12% amalgam

$$= 12/88 \times 2 = 0.273 \text{ g}$$



$$1 \text{ mole } 2 \times 96500\text{C}$$

$$112.40\text{g}$$

Charge required to deposit 0.273 g of Cd

$$= 2 \times 96500 / 112.40 \times 0.273 \text{ coulomb}$$

Charge = ampere \times second

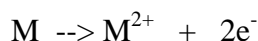
$$\text{Second} = 2 \times 96500 \times 0.273 / 112.40 \times 5 = 93.75$$

Example 23: Assume that impure copper contains irons, gold and silver as impurities. After passing a current of 140 ampere for 482. Second, the mass of anode decreased by 22.260 g and the cathode increased in mass by 22.011 g. Estimate the percentage of iron and copper originally present.

Solution: The increase at the cathode is due to copper only. Hence, there is 22.011 g of copper and rest impurities of iron, gold and silver.

$$\text{Mass of impurities} = (22.260 - 22.011) = 0.249 \text{ g}$$

At anode, only copper and iron are oxidised; the gold and silver collect below anode in the form of anodic mud.



(copper and iron)

$$\text{No. of moles of metal oxidised} = (140 \times 482.5) / (2 \times 96500) = 0.35$$

$$\text{No. of moles of copper} = 22.011 / 63.5 = 0.3466$$

$$\text{No. of moles of iron} = (0.35 - 0.3466) = 0.0034$$

Some Solved Examples

Example 24: 1.0 N solution of a salt surrounding two platinum electrodes 2.1 cm apart and 4.2 sq.cm in area was found to offer a resistance of 50 ohm. Calculate the equivalent conductivity of the solution.

Solution: Given, $l = 2.1 \text{ cm}$, $a = 4.2 \text{ sq cm}$, $R = 50 \text{ ohm}$

$$\text{Specific conductance, } k = l/a \cdot 1/R$$

$$\text{or } k = 2.1 / 4.2 \times 1 / 50 = 0.01 \text{ ohm}^{-1} \text{ cm}^{-1}$$

$$\text{Equivalent conductivity} = k \cdot V$$

$$V = \text{the volume containing 1 g equivalent} = 1000 \text{ mL}$$

$$\begin{aligned}\text{So Equivalent conductivity} &= 0.01 \times 1000 \\ &= 10 \text{ ohm}^{-1} \text{ cm}^{-2} \text{ equiv}^{-1}\end{aligned}$$

Example 25: Specific conductance of a decinormal solution of KCl is $0.0112 \text{ ohm}^{-1} \text{ cm}^{-1}$. The resistance of a cell containing the solution was found to be 56. What is the cell constant?

Solution: We know that

$$\text{Sp. conductance} = \text{Cell constant} \times \text{conductance}$$

$$\text{or Cell constant} = (\text{Sp. conductance}) / \text{Conductance}$$

$$= \text{Sp. conductance} \times \text{Resistance}$$

$$= 0.0112 \times 56$$

$$= 0.06272 \text{ cm}^{-1}$$

Example 26: The specific conductivity of 0.02 M KCl solution at 25°C is $2.768 \times 10^{-3} \text{ ohm}^{-1} \text{ cm}^{-1}$. The resistance of this solution at 25°C when measured with a particular cell was 250.2 ohms. The resistance of 0.01 M CuSO_4 solution at 25°C measured with the same cell was 8331 ohms. Calculate the molar conductivity of the copper sulphate solution.

Solution: Cell constant = (Sp.cond.of KCl)/(Conductance of KCl)

$$= (2.768 \times 10^{-3}) / (1/250.2)$$

$$= 2.768 \times 10^{-3} \times 250.2$$

For 0.01 M CuSO_4 solution

$$\text{Sp. conductivity} = \text{Cell constant} \times \text{conductance}$$

$$= 2.768 \times 10^{-3} \times 250.2 \times 1/8331$$

$$\text{Molar conductance} = \text{Sp. cond.} \times 1000/c$$

$$= (2.768 \times 10^{-3} \times 25.2) / 8331 \times 1000 / (1/100)$$

Example 27: The equivalent conductances of sodium chloride, hydrochloric acid and sodium acetate at infinite dilution are 126.45, 426.16 and 91.0 $\text{ohm}^{-1}\text{cm}^2 \text{equiv}^{-1}$, respectively, at 25°C. Calculate the equivalent conductance of acetic acid at infinite dilution.

Solution: According to Kohlrausch's law, $\Lambda_{\infty}(\text{CH}_3\text{COONa}) = \lambda_{\infty}(\text{CH}_3\text{COO}^-) + \lambda_{\infty}(\text{Na}^+) = 91.0 \dots\dots(i)$

$$\Lambda_{\infty}\text{HCl} = \lambda_{\text{H}^+} + \lambda_{\text{Cl}^-} = 426.16 \dots(ii)$$

$$\Lambda_{\infty}\text{NaCl} = \lambda_{\text{Na}^+} + \lambda_{\text{Cl}^-} = 126.45 \dots(iii)$$

Adding Eqs. (i) and (ii) and subtracting (iii),

$$\lambda_{\text{CH}_3\text{COO}^-} + \lambda_{\text{Na}^+} + \lambda_{\text{H}^+} + \lambda_{\text{Cl}^-} - \lambda_{\text{Na}^+} + \lambda_{\text{Cl}^-}$$

$$= 91.0 + 426.16 - 126.45$$

$$\lambda_{\text{CH}_3\text{COO}^-} + \lambda_{\text{H}^+} = \Lambda_{\infty}\text{CH}_3\text{COOH} = 390.7 \text{ ohm}^{-1} \text{ cm}^2 \text{ equiv}^{-1}$$

Example 28: The equivalent conductivity of N/10 solution of acetic acid at 25°C is 14.3 $\text{ohm}^{-1} \text{cm}^2 \text{equiv}^{-1}$. Calculate the degree of dissociation of CH_3COOH if Λ_{∞} is 390.71.

Solution: $\Lambda_{\infty}\text{CH}_3\text{COOH} = 390.7 \text{ ohm}^{-1} \text{ cm}^2 \text{ equiv}^{-1}$

$$\Lambda_{\infty}\text{CH}_3\text{COOH} = 143.3 \text{ ohm}^{-1} \text{ cm}^2 \text{ equiv}^{-1}$$

$$\text{Degree of dissociation, } \alpha = \Lambda_v / \Lambda_{\infty} = 14.3 / 390.71$$

$$= 0.0366 \text{ i.e., } 3.66\% \text{ dissociated}$$

Example 29: A decinormal solution of NaCl has specific conductivity equal to 0.0092. If ionic conductances of Na^+ and Cl^- ions at the same temperature are 43.0 and 65.0 ohm^{-1} respectively, calculate the degree of dissociation of NaCl solution.

Solution: Equivalent conductance of N/10 NaCl solution

$$\Lambda_v = \text{Sp. conductivity} \times \text{dilution}$$

$$= 0.0092 \times 10,000$$

$$= 92 \text{ ohm}^{-1}$$

$$\Lambda_\infty = \lambda_{\text{Na}^+} + \lambda_{\text{Cl}^-}$$

$$= 43.0 + 65.0$$

$$= 108 \text{ ohm}^{-1}$$

Degree of dissociation,