Solved Examples On Electrochemistry

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

Solution: Charge on one ion of N\(^{3-}\)

\[
= 3 \times 1.6 \times 10^{-19} \text{ coulomb}
\]

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

\[
= 3 \times 1.6 \times 6.02 \times 10^{23}
\]

\[
= 2.89 \times 10^5 \text{ coulomb}
\]

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

Solution: (a) The reduction reaction is

\[
\text{Al}^{3+} + 3e^- \rightarrow \text{Al}
\]

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

\[
\text{Mn}_4^- + 8H^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4H_2O
\]

1 mole 5 mole

\[
Q = 5 \times F
\]

\[
= 5 \times 96500 = 48500 \text{ coulomb}
\]
Example 3. How much electric charge is required to oxidise (a) 1 mole of H$_2$O to O$_2$ and (b) 1 mole of FeO to Fe$_2$O$_3$?

Solution: (a) The oxidation reaction is

\[ \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2 + 2\text{H}^+ + 2\text{e}^- \]

1 mole \hspace{1cm} 2 mole

\[ Q = 2 \times F \]

\[ = 2 \times 96500 = 193000 \text{ coulomb} \]

(b) The oxidation reaction is

\[ \text{FeO} + \frac{1}{2} \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{Fe}_2\text{O}_3 + \text{H}^+ + \text{e}^- \]

\[ Q = F = 96500 \text{ coulomb} \]

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.

\[ \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag} \]

1 mole \hspace{1cm} 1 mole

\[ 108 \text{ g} \hspace{1cm} 1 \text{ F} \]

\[ \text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} \]

1 mole \hspace{1cm} 2 mole

\[ 63.5 \text{ g} \hspace{1cm} 2 \text{ F} \]

and \[ \text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe} \]

1 mole \hspace{1cm} 3 mole
Hence,  
\[
\begin{align*}
56 \text{ g} & \quad 3 \text{ F} \\
\text{Ag deposited} &= 108 \times 0.4 = 43.2 \text{ g} \\
\text{Cu deposited} &= 63.5/2 \times 0.4 = 12.7 \text{ g} \\
\text{and} & \quad \text{Fe deposited} = 56/3 \times 0.4 = 7.47 \text{ g}
\end{align*}
\]

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

\[
2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-
\]

\[
\begin{align*}
71.0 \text{ g} & \quad 71.0 \text{ g} \\
2 \times 96500 \text{ coulomb} & \quad 1 \text{ mole}
\end{align*}
\]

\[
Q = I \times t = 100 \times 5 \times 600 \text{ coulomb}
\]

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

\[
= \frac{1}{2 \times 96500} \times 100 \times 5 \times 60 \times 60 = 9.3264 \text{ mole}
\]

Volume of \text{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

\[
\text{Watt} = \text{ampere} \times \text{volt}
\]

\[
100 = \text{ampere} \times 110
\]

\[
\text{Ampere} = \frac{100}{110}
\]

\[
\text{Quantity of charge} = \text{ampere} \times \text{second}
\]
= 100/110×10×60×60 coulomb

The cathodic reaction is

\[ \text{Cd}^{2+} + 2e^- \rightarrow \text{Cd} \]

\[ 112.4 \text{ g} \times 2 \times 96500 \text{ C} \]

Mass of cadmium deposited by passing \( 100/110×10×60×60 \)

Coulomb charge = \( 112.4/(2\times96500)\times100/110\times10\times60\times60 = 19.0598 \text{ g} \)

**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

\[ \frac{\text{Mass of Au deposited}}{\text{Mass of Cu deposited}} = \frac{\text{Eq. mass of Au}}{\text{Eq. Mass of Cu}} \]

Eq. mass of Au = 197/3; 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.

\[ E = Z \times 96500 \]

or \[ Z = E/96500 = 63.5/(2\times96500) \]

Applying \( W = Z \times I \times t \)

\[ T = 5 \text{ hour} = 5 \times 3600 \text{ second} \]

\[ 4.7625 = 63.5/(2\times96500) \times 1 \times 5 \times 3600 \]

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

\[ \text{Mass of silver} = \text{Volume} \times \text{density} \]
\[ \text{Mass of silver} = \text{Area} \times \text{thickness} \times \text{density} \]

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

Mass of silver to be deposited = 80 \times 0.0005 \times 10.5
\[ = 0.42 \text{ g} \]

Applying to silver \( E = Z \times 96500 \)
\[ Z = \frac{108}{96500} \text{ g} \]

Let the current be passed for \( r \) seconds.

We know that

\[ W = Z \times I \times t \]

So, 0.42 = \( \frac{108}{96500} \times 3 \times t \)

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

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 = \( Z \times 96500 \)

or \( Z = \frac{35.5}{96500} \text{ g} \)

Now, applying the formula
W = Z × I × t

Where W = 10 g, \( Z = \frac{35.5}{96500} \) t=60×60=3600 second

\[ I = \frac{10\times96500}{35.5\times96500} = 7.55 \text{ ampere} \]

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 coulomb deposit copper = 0.2964 g

96500 coulomb deposit copper = \( \frac{0.2964}{900} \times 96500 = 31.78 \) g

Thus, 31.78 is the equivalent mass of copper.

At. mass = Eq. mass × 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

\[ 2\text{SnCl}_2 \rightarrow \text{SnCl}_4 + \text{Sn} \]

\[ 2\times190 \text{ g} \quad 261 \text{ g} \quad 119 \text{ g} \]

119 g of Sn is deposited by the decomposition of 380 g of SnCl$_2$

So, 0.119 g of SnCl$_2$ of Sn is deposited by the decomposition of

\[ \frac{380}{119\times0.119} = 0.380 \text{ g of SnCl}_2 \]

Remaining amount of SnCl$_2$ = (19-0.380) = 18.62 g
380 g of SnCl$_2$ produce = 261 g of SnCl$_4$

So 0.380 g of SnCl$_2$ produce = $\frac{261}{380} \times 0.380 = 0.261$ g of SnCl

Thus, the ratio SnCl$_2$ : SnCl$_4$ = $\frac{18.2}{0.261}$, 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:

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \text{ (Cathode)} \]

1 mole \hspace{0.5cm} 2 \times 96500 \text{ C} \hspace{0.5cm} \text{Cu} \rightarrow \text{Cu}^{2+} + 2e^- \text{ (Anode)} \]

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 = $\frac{63.5}{(2 \times 96500)} \times 2.68 \times 60 \times 60$

= 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:

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

1 mole \hspace{0.5cm} 2 \times 96500 \text{ C} \hspace{0.5cm} 63.5 \text{ g of copper deposited by passing charge} = 2 \times 96500 \text{ 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

\[Q = I \times t \]

\[1930 = I \times 60 \times 60 \]

\[I = 1930 / 3600 = 0.536 \text{ ampere} \]

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 Ni(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

\[\text{Ni}^{2+} + 2e^- \rightarrow \text{Ni} \]

1 mole 2 \times 96500 C

Quantity of electric charge passed

\[= 3.7 \times 6 \times 60 \times 60 \text{ coulomb} = 79920 \text{ coulomb} \]

Number of moles of Ni(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 Cu\(^{2+}\) = 0.4/31.75 = 0.0126 g equivalent

At the same time, the oxygen deposited at anode

= 0.0126 g equivalent

= 8/32 × 0.0126 = 0.00315 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 × 7 × 60 = 504 coulomb

So, Oxygen liberated = 1/96500 × 504 = 0.00523 g equivalent

= 8/32 × 0.00523 = 0.001307 g mole

Hydrogen liberated = 0.00523 g equivalent

= 1/2 × 0.00523 = 0.00261 g mole

Total gases evolved = (0.00315 + 0.001307 + 0.0026 1) g mole

= 0.007067 g mole

Volume of gases evolved at NTP

= 22400 × 0.007067 mL

= 158.3 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

Amount of actual charge passed = \(90/100 \times 1.70 \times 230\)

= 351.9 coulomb

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

= \([0.160-(0.000182 \times 1000)/300]M\)

= 0.154 M

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

\[\text{C}_6\text{H}_5\text{NO}_2 + 3\text{H}_2 \rightarrow \text{C}_6\text{H}_5\text{NH}_2 + 2\text{H}_2\text{O}\]

\[\begin{align*}
123 \text{ g} & \quad 6\text{ g} \\
1 \text{ mole} & \quad 3 \text{ mole}
\end{align*}\]

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%

= \(2 \times 57900 = 115800\) coulomb

Energy consumed = \(115800 \times 3.0 = 347400 \text{ J}\)

= 347.4 \text{ kJ}
Example 18: An aqueous solution of sodium chloride on electrolysis gives H₂(g), Cl₂(g) and NaOH according to the reaction:

\[ 2\text{Cl}^- (aq) + 2\text{H}_2\text{O} \rightarrow 2\text{OH}^- (aq) + \text{H}_2(g) + \text{Cl}_2(g) \]

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 Cl₂? 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

2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^- \text{ (at anode)}

2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^- \text{ (at cathode)}

1 kg of Cl₂ = 1000/71.0 = 14.08 mole

Charge to produce one mole of Cl₂ = 2 \times 96500 coulomb

Charge to produce 14.08 mole of Cl₂ = 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 second = 48.699 hour

\text{OH}^- \text{ ions produced} = 2 \times \text{mole of Cl₂}

= 2 \times 14.08 = 28.16 mole

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

Example 19: Chromium metal can be plated out from an acidic solution containing \text{CrO}_3 according to the following reaction:

\[ \text{CrO}_3 + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{Cr} + 3\text{H}_2\text{O} \]

Solution: \[ \text{CrO}_3 + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{Cr} + 3\text{H}_2\text{O} \]
6 × 96500C 1 mole

52 g.

Mass of chromium plated out by 24000 coulomb charge

= 52/6×96500 × 24000 = 2.155 g

Charge required for plating out 1.5 g of chromium

= 6×96500/52 × 1.5 = 16701.92 coulomb

Time = charge/current = 1336.15 second

= 22.27 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 = 40/1000 × 600 = 24 g

Amount of NaOH formed = 40/1000 × 600 = 24 g

31.75 g of Cu = 1 g equivalent of Cu.

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

1 g equivalent of NaOH = 40 g

% yield = 24/40 × 100 = 60

**Example 21:** Peroxy disulphuric acid (H$_2$S$_2$O$_8$) can be prepared by electrolytic oxidation of H$_2$SO$_4$ as:

2H$_2$SO$_4$ --> H$_2$S$_2$O$_8$ + 2H$^+$ +2e$^-$

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:

Anodic reaction $2\text{H}_2\text{SO}_4 \rightarrow \text{H}_2\text{S}_2\text{O}_8 + 2\text{H}^+ + 2\text{e}^-$

$2\text{H}_2\text{O} \rightarrow 4\text{H}^+ + \text{O}_2 + 2\text{e}^-$

Cathodic reaction $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{OH}^- + \text{H}_2$

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

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

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

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

$= 0.448$

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 $\text{CdCl}_2$ using a mercury cathode. Find how long a current of 5 amperes 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:** $2 \text{ g } \text{Hg require Cd to prepare 12% amalgam}$

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

$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$

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

$112.40\text{g}$

Charge required to deposit 0.273 g of Cd

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

Charge = ampere $\times$ second

Second $= 2*96500*0.273/112.40*5 = 93.75$
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.

Mass of impurities = (22.260 - 22.011) = 0.249 g

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

\[ M \rightarrow M^{2+} + 2e^- \]

(copper and iron)

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

No. of moles of copper = 22.011/63.5 = 0.3466

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, \(l = 2.1\) cm, \(a = 4.2\) sq cm, \(R = 5\) ohm

Specific conductance, \(k = l/a.1/R\)

or \(k = 2.1/4.2\times1/50 = 0.01\) ohm\(^{-1}\) cm\(^{-1}\)

Equivalent conductivity = \(k = V\)

\(V = \) the volume containing 1 g equivalent = 1000 mL
So Equivalent conductivity = 0.01 × 1000

= 10 ohm\(^{-1}\) cm\(^{-2}\) equiv\(^{-1}\)

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

**Solution:** We know that

Sp. conductance = Cell constant × conductance

or Cell constant = (Sp.conductance)/Conductance

= Sp. conductance × Resistance

= 0.0112 × 56

= 0.06272 cm\(^{-1}\)

**Example 26:** The specific conductivity of 0.02 M KCl solution at 25\(^0\)C is 2.768 × 10\(^{-3}\) ohm\(^{-1}\) cm\(^{-1}\). The resistance of this solution at 25\(^0\)C when measured with a particular cell was 250.2 ohms. The resistance of 0.01 M CuSO\(_4\) solution at 25\(^0\)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)

= (2.768×10\(^{-3}\))/250.2

= 2.768 × 10\(^{-3}\) × 250.2

For 0.01 M CuSO\(_4\) solution

Sp. conductivity = Cell constant × conductance

= 2.768 × 10\(^{-3}\) × 250.2 × 1/8331

Molar conductivity = Sp. cond. × 1000/c

= (2.768×10\(^{-3}\)×25.2)/8331×1000/(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 ohm⁻¹cm²equiv⁻¹, respectively, at 25°C. Calculate the equivalent conductance of acetic acid at infinite dilution.

Solution: According to Kohlrausch's law, \( \lambda_{(\infty CH_3 COONa)} = \lambda_{(CH_3 COO^-)} + \lambda_{(Na^+)} = 91.0 \) ....(i)
\[ \lambda_{\infty HCl} = \lambda_{H^+} + \lambda_{Cl^-} = 426.16 \] ...(ii)
\[ \lambda_{\infty NaCl} = \lambda_{Na^+} + \lambda_{Cl^-} = 126.45 \] ...(iii)

Adding Eqs. (i) and (ii) and subtracting (iii),
\[ \lambda_{CH_3 COO^-} + \lambda_{Na^+} + \lambda_{H^+} + \lambda_{Cl^-} - \lambda_{Na^+} + \lambda_{Cl^-} = 91.0 + 426.16 - 126.45 \]
\[ \lambda_{CH_3 COO^-} + \lambda_{H^+} = \lambda_{\infty CH_3 COOH} = 390.7 \text{ ohm}^{-1} \text{ cm}^2 \text{ equiv}^{-1} \]

Example 28: The equivalent conductivity of N/10 solution of acetic acid at 25°C is 14.3 ohm⁻¹cm²equiv⁻¹. Calculate the degree of dissociation of CH₃COOH if \( \Lambda_{\infty} \) is 390.71.

Solution: \( \Lambda_{\infty CH_3 COOH} = 390.7 \text{ ohm}^{-1} \text{ cm}^{-2} \text{ equiv}^{-1} \)
\[ \Lambda_{(\infty CH_3 COOH)} = 143.3 \text{ ohm}^{-1} \text{ cm}^{-2} \text{ equiv}^{-1} \]

Degree of dissociation, \( \alpha = \Lambda_{\infty} / \Lambda_{\infty} = 14.3/390.71 \)
\[ = 0.0366 \text{ i.e., 3.66% dissociated} \]

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
\[ \Lambda_v = \text{Sp. conductivity} \times \text{dilution} \]

\[ = 0.0092 \times 10,000 \]

\[ = 92 \text{ ohm}^{-1} \]

\[ \Lambda \infty = \lambda_{Na^+} + \lambda_{CI^-} \]

\[ = 43.0 + 65.0 \]

\[ = 108 \text{ ohm}^{-1} \]

Degree of dissociation,