Chapter.3
Electrochemistry
Class-XII
Subject-Chemistry

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

Answer 3.1

The order in which metal displaces each other from the solution of their salts is:
Mg, Al, Zn, Fe, Cu

3.2 Given the standard electrode potentials,

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

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

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

Arrange these metals in their increasing order of reducing power.

Answer 3.2

Lower the reduction potential, higher will be reducing power.

Therefore, the order of reducing power of the given metals increases as:

Ag < Hg < Cr < Mg < K

3.3 Depict the galvanic cell in which the reaction

$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ **takes place. Further show:**

i. **Which of the electrode is negatively charged?**
ii. **The carriers of the current in the cell.**
iii. **Individual reaction at each electrode.**

Answer 3.3

The galvanic cell of the given reaction can be represented as:

$$Zn_{(s)} \left| Zn^{2+}_{(aq)} \, \mathrm{P} \, Ag^+_{(aq)} \right| Ag_{(s)}$$
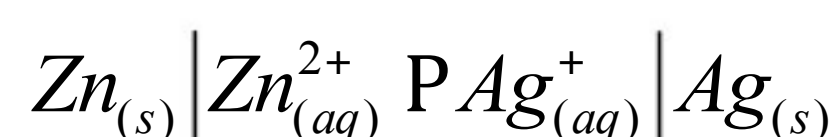
i. Zn electrode is negatively charged.
ii. Ions are carriers of current in the cell. Current will flow from silver to zinc in the external circuit.
iii. The reaction takes place at the anode is:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$
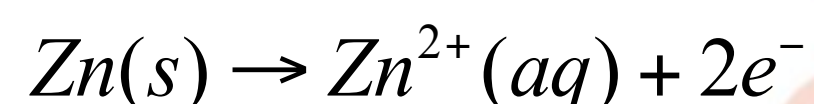
The reaction takes place at the cathode is: $Ag^+(aq) + e^- \rightarrow Ag(s)$

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

i. $2Cr(s) + 3Cd^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cd$
ii. $Fe^{2+}(aq) + Ag^+(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$

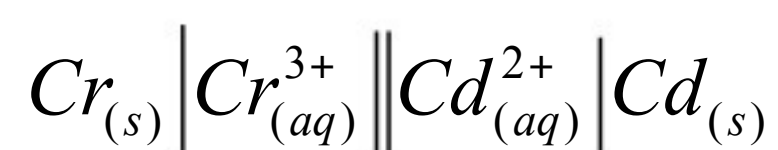**Calculate the $\Delta_r G^\theta$ and equilibrium constant of the reactions.**

Answer 3.4

i.

$E^\Theta_{Cr^{3+}/Cr} = 0.74 \ V$

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

Galvanic cell of given reaction is:

$Cr_{(s)} \left| Cr^{3+}_{(aq)} \right\| Cd^{2+}_{(aq)} \left| Cd_{(s)} \right.$

Now,

$$E_{cell}^{\Theta} = E_R^{\Theta} - E_L^{\Theta}$$
$$= -0.40 - (-0.74)$$
$$= 0.34 \ V$$
$$\Delta_r G^{\Theta} = -nFE_{cell}^{\Theta}$$

We know that,

$n = 6$

$$\Delta_r G^{\Theta} = -nFE_{cell}^{\Theta}$$
$$= -6 \times 96487 \times 0.34$$
$$= -196.83 \ Kj / mol$$

Also,

$$\Delta_r G^{\Theta} = -2.303 RT \log K$$

$$\log K = -\frac{\Delta_r G^{\Theta}}{2.303 RT}$$
$$= -\frac{196.83 \times 1000}{2.303 \times 8.314 \times 298} = 34.496$$

Thus,

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

$K = 3.13 \times 10^{34}$
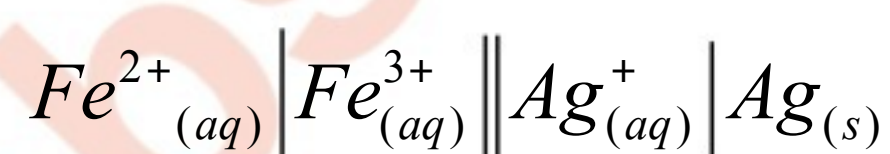
ii.

$$E_{Fe^{3+}/Fe^{2+}}^{\Theta} = 0.77 \ V$$
$$E_{Ag^+/Ag}^{\Theta} = 0.80 \ V$$

Galvanic cell of given reaction is:

$$Fe^{2+}_{(aq)} \Big| Fe^{3+}_{(aq)} \Big\| Ag^+_{(aq)} \Big| Ag_{(s)}$$

Now,

$$E_{cell}^{\Theta} = E_R^{\Theta} - E_L^{\Theta}$$
$$= 0.80 - 0.77$$
$$= 0.03 \ V$$

We know that,

$n = 1$

$\Delta_r G^\Theta = -nFE^\Theta_{cell}$

$= -1 \times 96487 \times 0.03$

$= -2.89 \ Kj/mol$

Also,

$\Delta_r G^\Theta = -2.303 RT \log K$

$\log K = -\dfrac{\Delta_r G^\Theta}{2.303 RT}$

$\qquad = -\dfrac{2.89 \times 1000}{2.303 \times 8.314 \times 298} = 0.5073$

Thus,

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

$K = 3.2$

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

    i.   **Mg(s) | Mg^{2+} (0.001M) || Cu^{2+} (0.0001 M) | Cu(s)**
    ii.   **Fe(s) | Fe^{2+} (0.001M) || H$^+$ (1M)|H$_2$(g)(1bar) | Pt(s)**
    iii.   **Sn(s) | Sn^{2+} (0.050 M) || H$^+$ (0.020 M) | H$_2$ (g) (1 bar) | Pt(s)**
    iv.   **Pt(s) | Br$_2$ (l) | Br$^-$ (0.010 M) || H$^+$ (0.030 M) | H$_2$ (g) (1 bar) | Pt(s).**

Answer 3.5

    i.   Nernst equation for the given equation is:

$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log \frac{\left[Mg^{2+} \right]}{\left[Cu^{2+} \right]}$$

$$= \left[0.34 - (-2.36) \right] - \frac{0.0591}{2} \log \frac{0.001}{0.0001}$$

$$= 2.7 - \frac{0.0591}{2} \log 10$$

$$= 2.67 V$$

    ii.   Nernst equation for the given equation is:

$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log \frac{[Fe^{2+}]}{[H^+]^2}$$

$$= [0 - (-0.44)] - \frac{0.0591}{2} \log \frac{0.001}{1}$$

$$= 0.44 - 0.0295 \times (-3)$$

$$= 0.528V$$

iii. Nernst equation for the given equation is:

$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log \frac{[Sn^{2+}]}{[H^+]^2}$$

$$= [0 - (-0.14)] - \frac{0.0591}{2} \log \frac{0.050}{(0.02)^2}$$

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

$$= 0.14 - 0.062$$

$$= 0.078V$$

iv.

Nernst equation for the given equation is:
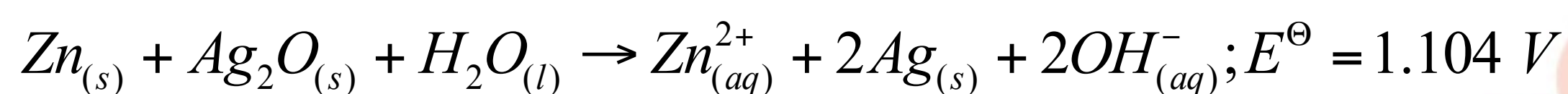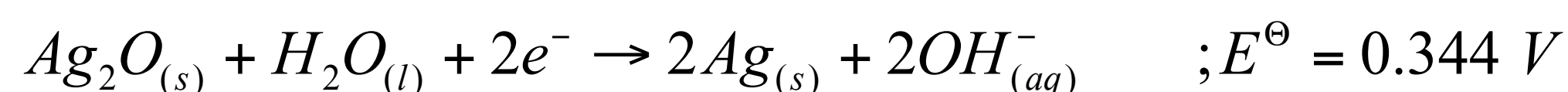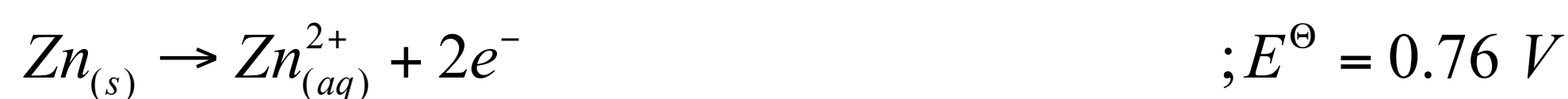
$$E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log \frac{1}{[Br^-]^2 [H^+]^2}$$

$$= [0 - 1.09] - \frac{0.0591}{2} \log \frac{1}{(0.01)^2 (0.03)^2}$$

$$= -1.09 - 0.0295 \times \log \left(\frac{1}{9 \times 10^{-8}} \right)$$

$$= -1.09 - 0.0295 \times 7.0453$$

$$= -1.298 \ V$$

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

$$Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + 2Ag(s) + 2OH^-(aq)$$

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

Answer 3.6

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

$$Ag_2O_{(s)} + H_2O_{(l)} + 2e^- \rightarrow 2Ag_{(s)} + 2OH^-_{(aq)} \qquad ; E^{\ominus} = 0.344 \; V$$

$$\overline{\rule{0pt}{0pt}\hspace{10cm}}$$

$$Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \rightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)} + 2OH^-_{(aq)} ; E^{\ominus} = 1.104 \; V$$

So, $E^{\ominus} = 1.104 \; V$

Also,

$$\Delta_r G^{\ominus} = -nFE^{\ominus}$$
$$= 2(96487)(1.04)$$
$$= -213.04 \; kj$$

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

Answer 3.7

Conductivity: - it is defined as the conductance of a solution of 1 cm in length and area of cross-section 1 sq. cm. The inverse of resistivity is called specific conductance or conductivity. It is represented by the symbol "κ".

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

Where ρ is resistivity

Also,

At any given concentration if a solution is kept between two platinum electrodes having unit area of cross-section & with a distance of unit length, then, conductivity is equal to the conductance (G).

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

As $a = 1$ & $l = 1$

Conductivity decreases with a decrease in concentration, for all the electrolytes weak and strong electrolytes. This happens because the number of ions that carry the current in a solution decreases with a decrease in concentration.

Molar conductivity: - at a given concentration, it can be defined as the conductance of volume V of a solution containing 1 mole of the electrolyte which is kept between two electrodes with the area A and distance of unit length.
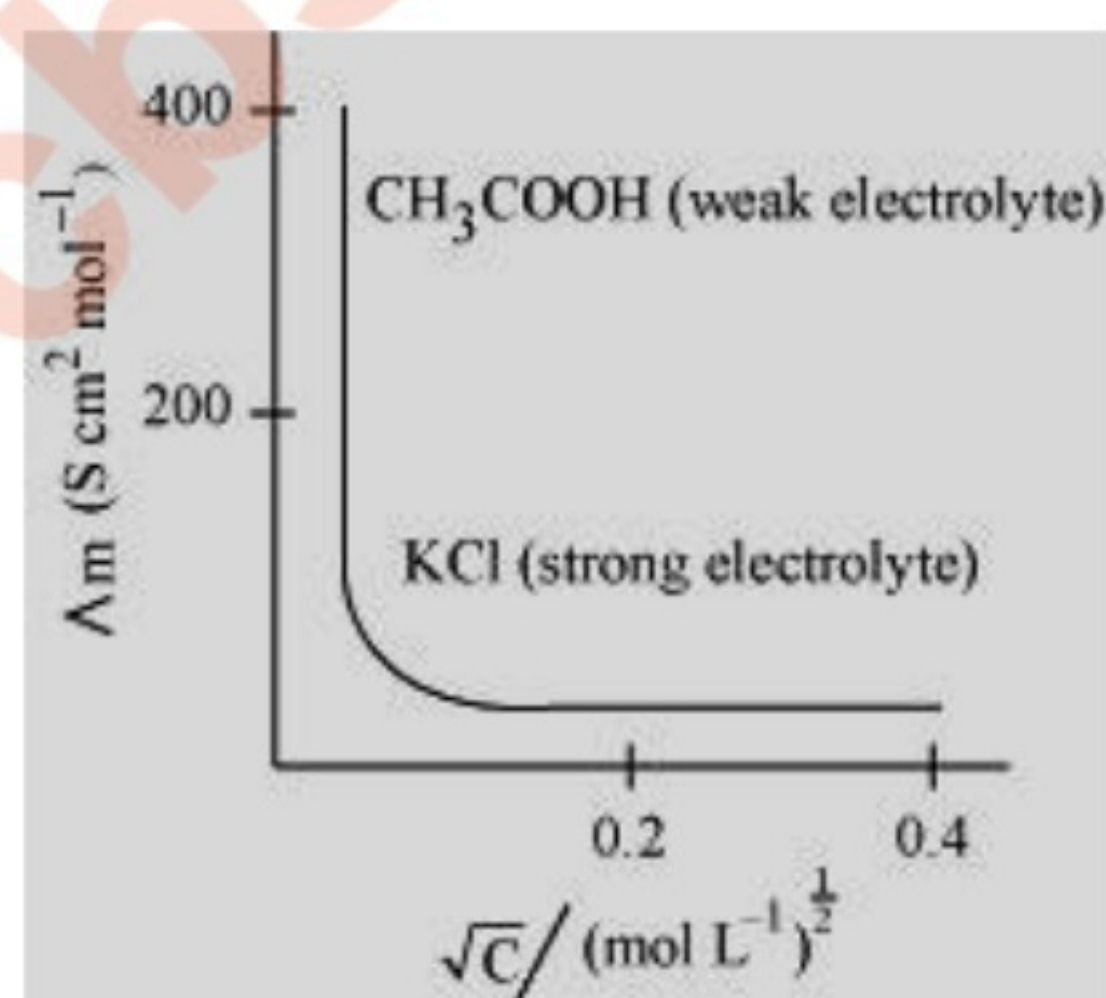
$$\Lambda_m = \kappa \frac{A}{l}$$
$$As\ l = 1,\ A = V$$
$$Hence,$$
$$\Lambda_m = \kappa V$$

It is inversely related to concentration thus it increases with a decrease in concentration. This happens because the total volume V of the solution containing one mole of the electrolyte increases on dilution.

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

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

Answer 3.8

κ = 0.0248 S cm^{-1}

c = 0.20 M

$$\Lambda_m = \frac{\kappa \times 1000}{c} = \frac{0.0248 \times 1000}{0.2} = 124 \ Scm^2 / mol$$

$\therefore$ Molar conductivity = 124 Scm^2 / mol

3.9 The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500 Ω. What is the cell constant if conductivity of 0.001M KCl solution at 298 K is 0.146 × 10^{-3} S cm^{-1}?

Answer 3.9

κ = 0.146 × 10^{-3} S cm^{-1}

R = 1500 Ω

$\therefore$ cell constant = $\kappa \times R$

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

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

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

Concentration/M	0.001	0.010	0.020	0.050	0.100
10^2 × κ/S m^{-1}	1.237	11.85	23.15	55.53	106.74

Calculate Λ_m for all concentrations and draw a plot between Λ_m and c½. Find the value of Λ_m^0.

Answer 3.10

1.

$\kappa = 1.237 \times 10^{-2}$ S m^{-1} = 1.237×10^{-4} S cm^{-1}

c = 0.001 M, $\sqrt{c}$ = 0.0316 M$^{1/2}$

$\Lambda_m = \dfrac{\kappa}{c} = \dfrac{1.237 \times 10^{-4} \times 1000}{0.001} = 123.7 \; Scm^2/mol$

2.

$\kappa = 11.85 \times 10^{-2}$ S m^{-1} = 11.85×10^{-4} S cm^{-1}

c = 0.010M, $\sqrt{c}$ = 0.1 M$^{1/2}$

$\Lambda_m = \dfrac{\kappa}{c} = \dfrac{11.85 \times 10^{-4} \times 1000}{0.010} = 118.5 \; Scm^2/mol$

3.

$\kappa = 23.15 \times 10^{-2}$ S m^{-1} = 23.15×10^{-4} S cm^{-1}

c = 0.020 M, $\sqrt{c}$ = 0.1414 M$^{1/2}$

$\Lambda_m = \dfrac{\kappa}{c} = \dfrac{23.15 \times 10^{-4} \times 1000}{0.020} = 115.8 \; Scm^2/mol$

4.

$\kappa = 55.53 \times 10^{-2}$ S m^{-1} = 55.53×10^{-4} S cm^{-1}

c = 0.050 M, $\sqrt{c}$ = 0.2236 M$^{1/2}$

$\Lambda_m = \dfrac{\kappa}{c} = \dfrac{55.53 \times 10^{-4} \times 1000}{0.050} = 111.11 \; Scm^2/mol$

5.

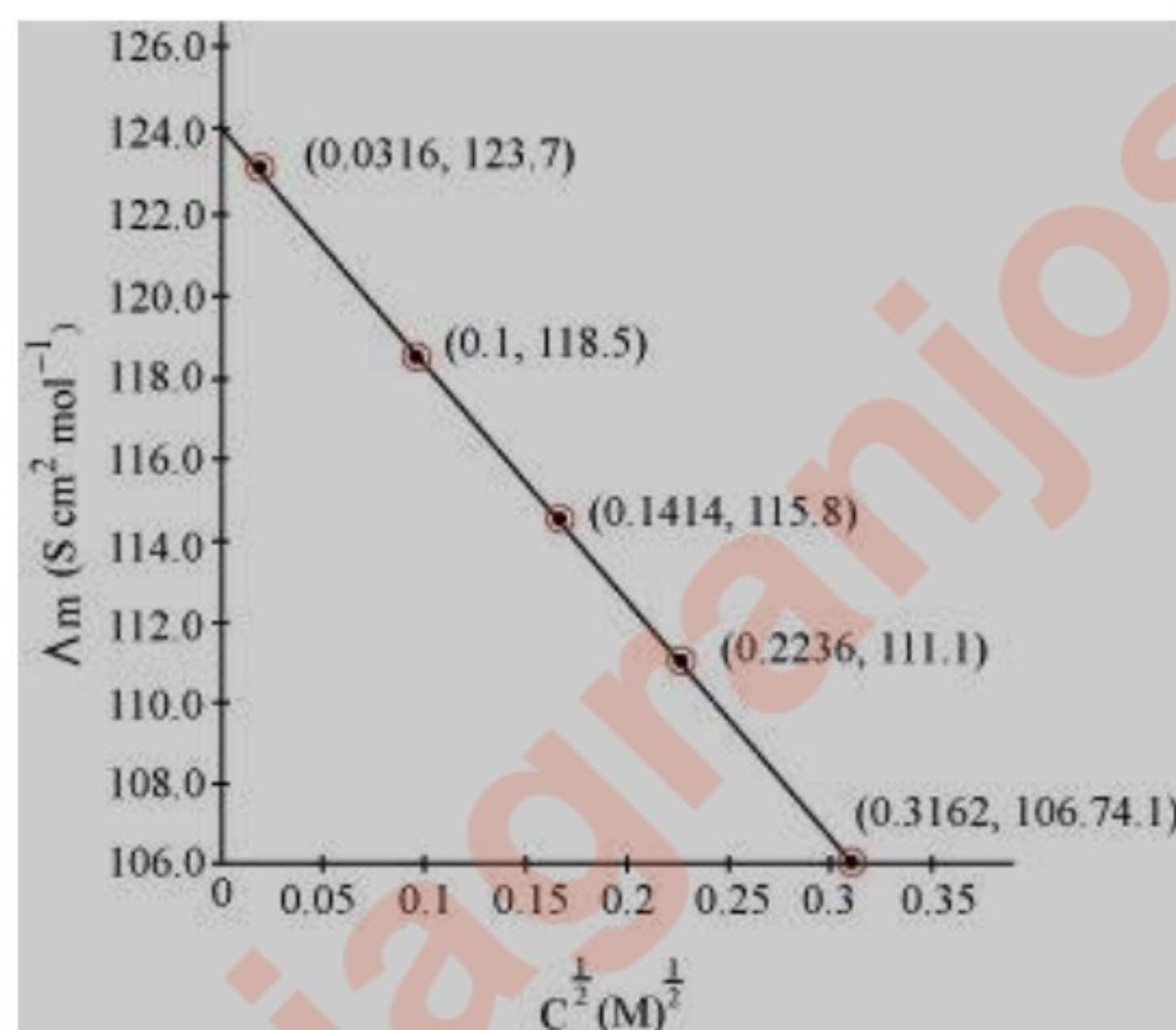κ = 106.74 × 10^{-2} S m^{-1} = 106.74 × 10^{-4} S cm^{-1}

c = 0.100 M, $\sqrt{c}$ = 0.3162 M$^{1/2}$

$$\Lambda_m = \frac{\kappa}{c} = \frac{106.74 \times 10^{-4} \times 1000}{0.100} = 106.74 \; Scm^2 / mol$$

Now, we have the following data:

$\sqrt{c}$	0.0316	0.1	0.1414	0.2236	0.3162
Λ_m	123.7	118.5	115.8	111.1	106.74

From the graph, $\Lambda_m^0 = 124 \; Scm^2 / mol$

3.11 Conductivity of 0.00241 M acetic acid is 7.896 × 10^{-5} S cm^{-1}. Calculate its molar conductivity and if Λ_m^0 for acetic acid is 390.5 S cm^{2}mol^{-1}, what is its dissociation constant?

Answer 3.11

κ = 7.896 × 10^{-5} S m^{-1}

c = 0.00241 mol L^{-1}

$$\Lambda_m = \frac{\kappa}{c} = \frac{7.896 \times 10^{-5} \times 1000}{0.00241} = 32.76 \ Scm^2 / mol$$

We have,

$$\Lambda_m^0 = 390.5 \ Scm^2 / mol$$

$$\alpha = \frac{\Lambda_m}{\Lambda_m^0} = \frac{32.76}{390.5} = 0.084$$

Dissociation constant, $\therefore K_a = \frac{c\alpha^2}{1-\alpha} = \frac{0.00241 \times 0.084 \times 0.084}{1 - 0.0084} = 1.86 \times 10^{-5} \ mol/l$

3.12 **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+}**

Answer 3.12

    i.   Charge = 3 F

            = 3 × 96487 C = 289461 C

    ii.   Charge = 2 F

            = 2 × 96487 C = 192974 C

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

        Here, +7 oxidation state is converted into +2
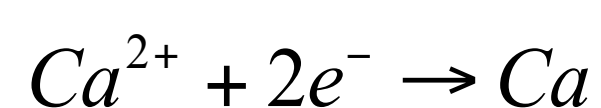
        Therefore, charge = 5 F

            = 5 × 96487 C = 482435 C

3.13 **How much electricity in terms of Faraday is required to produce?**

i. **20.0 g of Ca from molten CaCl$_2$**
ii. **40.0 g of Al from molten Al$_2$O$_3$**

Answer 3.13

i. As per the equation,

$$Ca^{2+} + 2e^- \rightarrow Ca$$
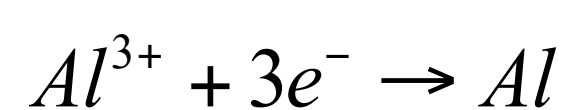
Electricity required to produce 40 g of calcium = 2 F

Hence, electricity required to produce 20 g of calcium = $\dfrac{2 \times 20}{40}$ F = 1 F

ii. As per the equation,

$$Al^{3+} + 3e^- \rightarrow Al$$
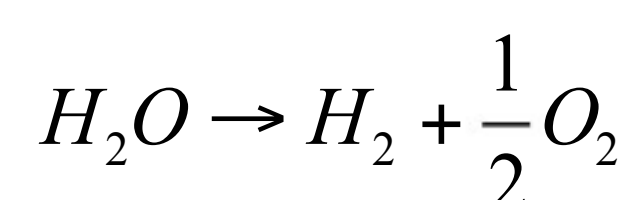
Electricity required to produce 27 g of Al = 3 F

Hence, electricity required to produce 40 g of Al = $\dfrac{3 \times 40}{27}$ = 4.44 F

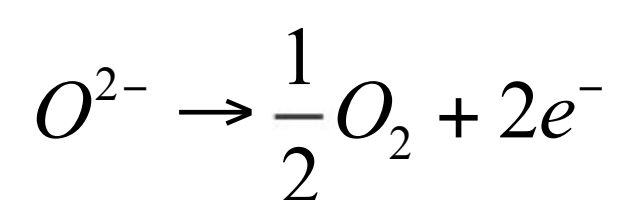**3.14 How much electricity is required in coulomb for the oxidation of?**

i. **1 mol of H$_2$O to O$_2$.**
ii. **1 mol of FeO to Fe$_2$O$_3$.**

Answer 3.14

i. As per the equation,

$$H_2O \rightarrow H_2 + \frac{1}{2}O_2$$

Now,

$$O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$$
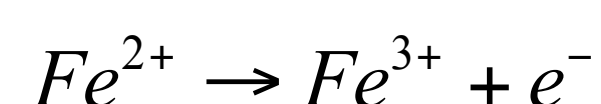
Electricity required for the oxidation of 1 mol of H_2O to O_2 = 2 F

= 2 × 96487 C = 192974 C

ii.   As per the equation,

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

Electricity required for the oxidation of 1 mol of FeO to Fe_2O_3 = 1 F = 96487 C

3.15 A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?
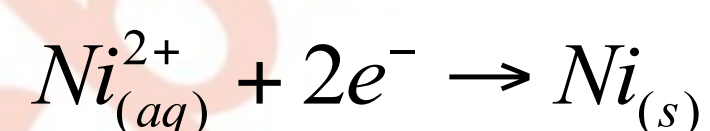
Answer 3.15

Current, I = 5A

Time, t = 20 × 60 = 1200 s

Charge = current × time

= 5 × 1200 = 6000 C

As per the equation,

$$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$$

Nickel deposited by 2 × 96487 C charge = 58.71 g

Hence, nickel deposited by 6000 C charge = $\dfrac{58.71 \times 6000}{2 \times 96487}$ g

= 1.825 g

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

3.16 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 B. How long did the current flow? What mass of copper and zinc were deposited?

Answer 3.16

As per the equation,

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$$
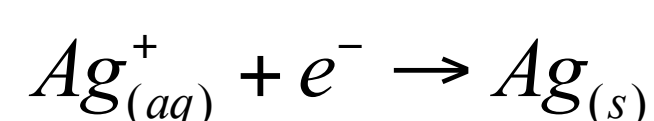
108 g of Ag is deposited by = 96487 C of charge

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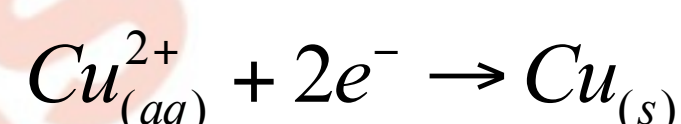
$$= 1295.43 \text{ C}$$

We have,

I = 1.5 A

$$Time = \dfrac{1295.43}{1.5} s = 863.6 \text{ s} = 864 \text{ s} = 14.40 \text{ min}$$

Now,

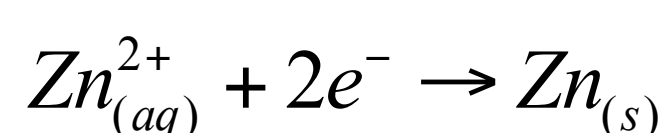$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

2 × 96487 C of charge deposit = 63.5 g of Cu

Then, 1295.43 C of charge will deposit = $\dfrac{63.5 \times 1295.43}{2 \times 96487}$ g

$$= 0.426 \text{ g of Cu}$$

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$$

2 × 96487 C of charge deposit = 65.4 g of Zn

Then, 1295.43 C of charge will deposit = $\dfrac{65.4 \times 1295.43}{2 \times 96487}$ g

= 0.439 g of Zn

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

  i. $Fe^{3+}_{(aq)}$ & $I^{-}_{(aq)}$

  ii. $Ag^{+}_{(aq)}$ & $Cu_{(s)}$

  iii. $Fe^{3+}_{(aq)}$ & $Br^{-}_{(aq)}$
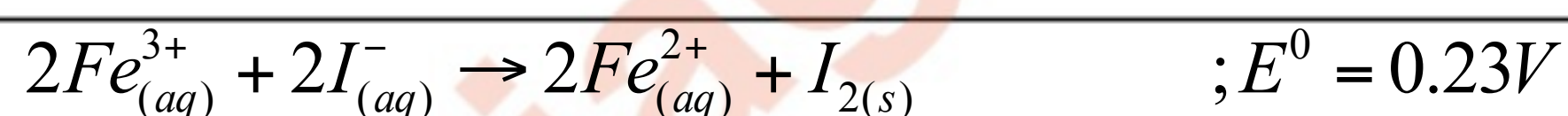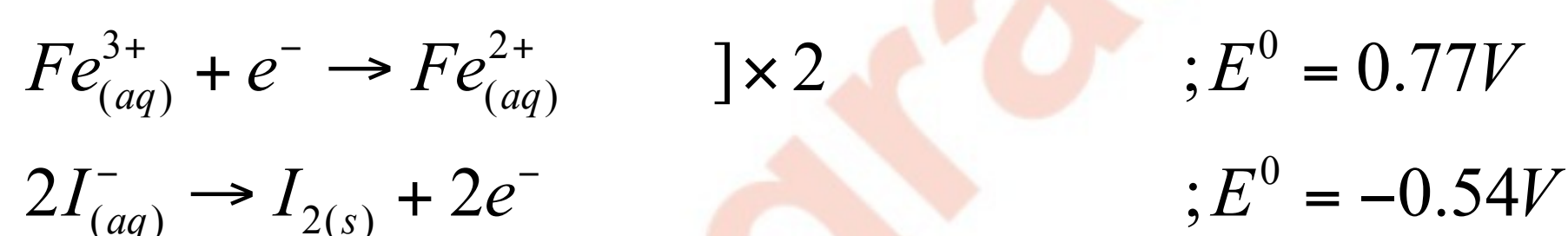
  iv. $Ag_{(s)}$ & $Fe^{3+}_{(aq)}$

  v. $Br_{2(aq)}$ & $Fe^{2+}_{(aq)}$

Answer 3.17

  i.

$Fe^{3+}_{(aq)} + e^{-} \rightarrow Fe^{2+}_{(aq)} \quad] \times 2 \qquad ; E^{0} = 0.77V$

$2I^{-}_{(aq)} \rightarrow I_{2(s)} + 2e^{-} \qquad\qquad ; E^{0} = -0.54V$

$\overline{}$

$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + I_{2(s)} \qquad ; E^{0} = 0.23V$

$E^{0} = positive$

$\therefore$ Reaction is fisible

  ii.

$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)} \quad] \times 2 \qquad ; E^{0} = 0.80 \ V$

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-} \qquad\qquad ; E^{0} = -0.34 \ V$

$\overline{}$

$2Ag^{+}_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)} \qquad ; E^{0} = 0.46 \ V$

$E^{0} = positive$

$\therefore$ Reaction is fisible

  iii.

$$\frac{Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)} \quad]\times 2 \qquad\qquad ; E^0 = 0.77V}{2Br^-_{(aq)} \rightarrow Br_{2(l)} + 2e^- \qquad\qquad ; E^0 = -1.09V}$$
$$2Fe^{3+}_{(aq)} + 2Br^-_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + Br_{2(l)} \qquad ; E^0 = -0.32V$$

$E^0 = negative$

∴ Reaction is not fisible

iv.

$$Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^- \qquad\qquad ; E^0 = -0.80 \ V$$

$$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)} \qquad\qquad ; E^0 = 0.77 \ V$$

$$Ag_{(s)} + Fe^{3+}_{(aq)} \rightarrow Ag^+_{(aq)} + Fe^{2+}_{(aq)} \qquad ; E^0 = -0.03 \ V$$

$E^0 = negative$

∴ Reaction is not fisible

v.

$$Br_{2(aq)} + 2e^- \rightarrow 2Br^-_{(aq)} \qquad\qquad ; E^0 = 1.09 \ V$$

$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^- \ ]\times 2 \qquad\qquad ; E^0 = -0.77 \ V$$

$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)} \qquad ; E^0 = 0.32 \ V$$
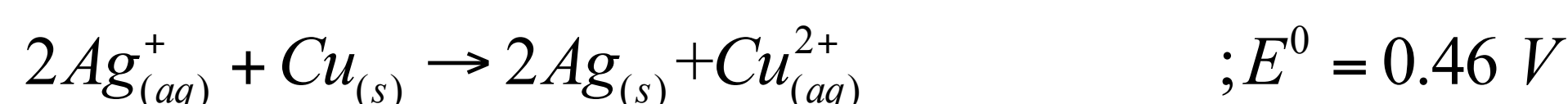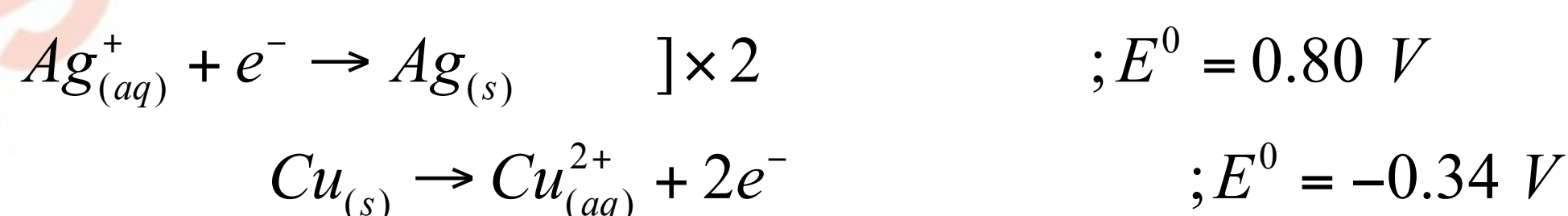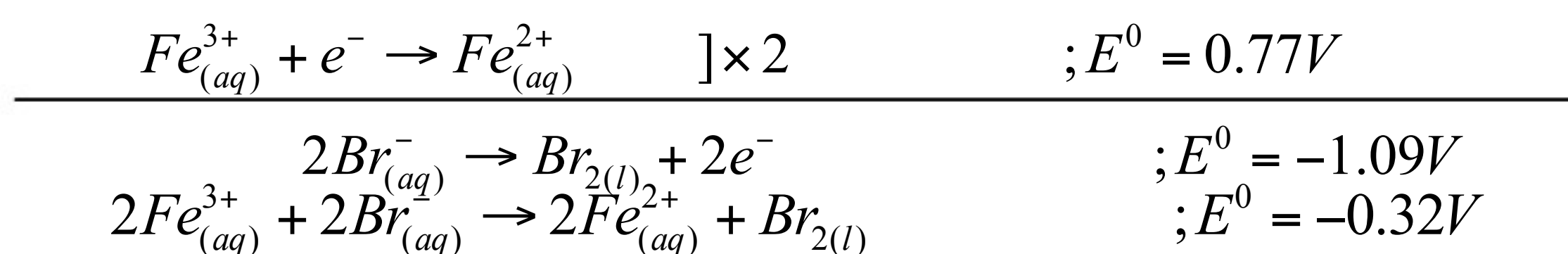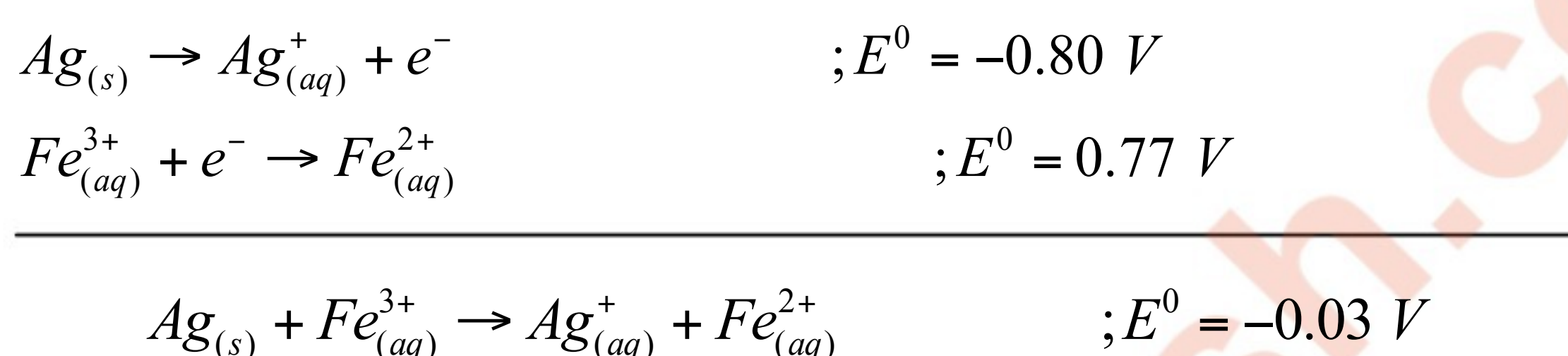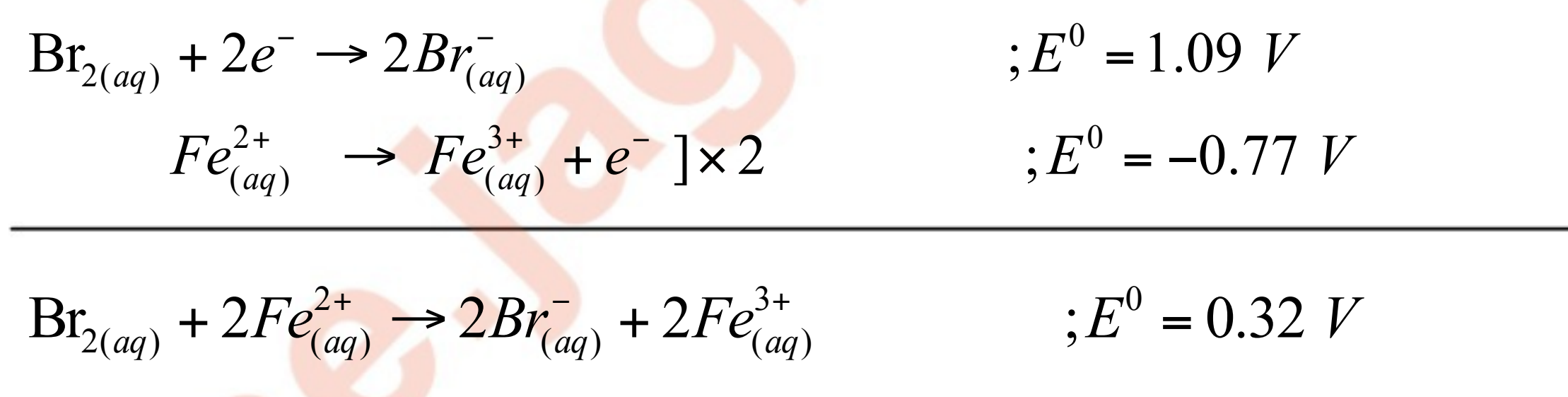
$E^0 = positive$

∴ Reaction is fisible

3.18 **Predict the products of electrolysis in each of the following:**

    i.   **An aqueous solution of AgNO₃ with silver electrodes.**
    ii.   **An aqueous solution of AgNO₃with platinum electrodes.**
    iii.   **A dilute solution of H₂SO₄with platinum electrodes.**
    iv.   **An aqueous solution of CuCl₂ with platinum electrodes.**

Answer 3.18

i. At cathode: reduction reaction
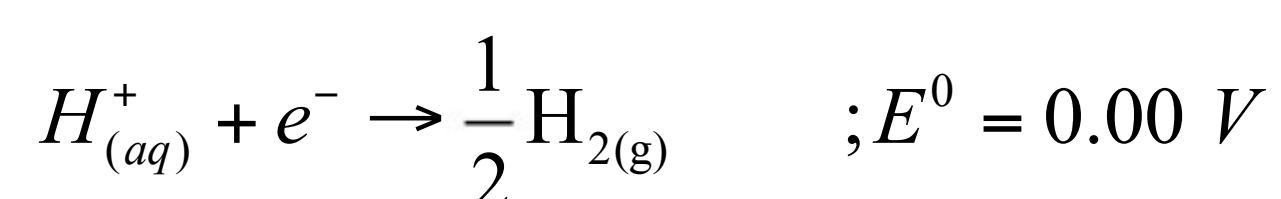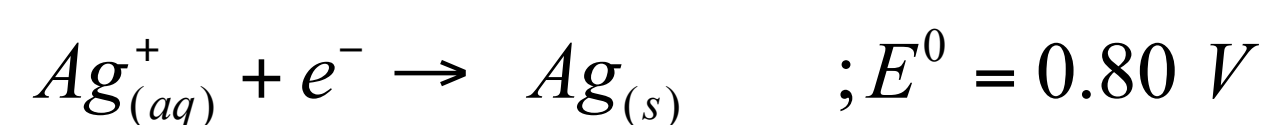
$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)} \qquad ; E^0 = 0.80 \ V$$

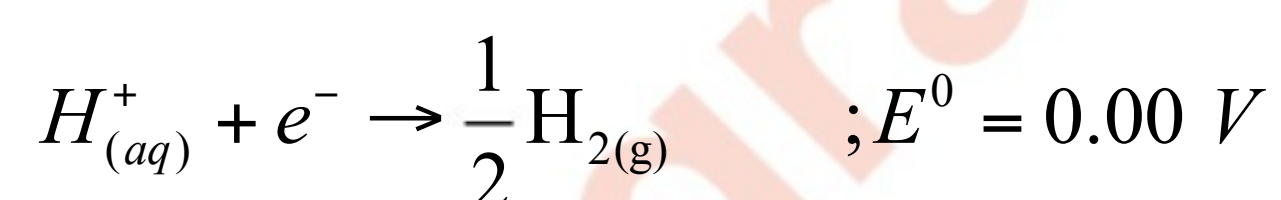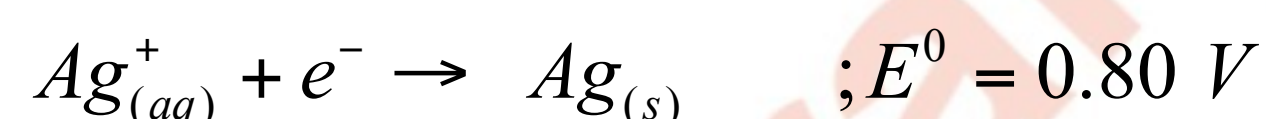$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2} H_{2(g)} \qquad ; E^0 = 0.00 \ V$$

As we know that the reaction which has higher E^0 value, it will takes place at cathode. Thus, deposition of silver will take place at the cathode.

At anode:

NO_3^- Ions will react with Ag anode. Thus, the silver electrode at the anode dissolves in the solution & form Ag^+.
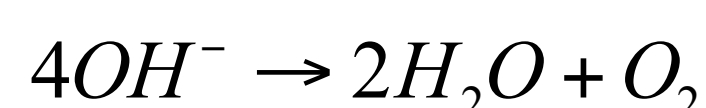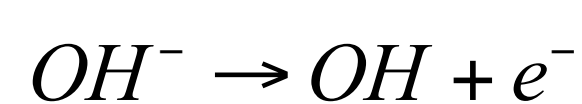
ii. At cathode: reduction reactions

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)} \qquad ; E^0 = 0.80 \ V$$

$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2} H_{2(g)} \qquad ; E^0 = 0.00 \ V$$

As we know that the reaction which has higher E^0 value, it will takes place at cathode. Thus, deposition of silver will take place at the cathode.
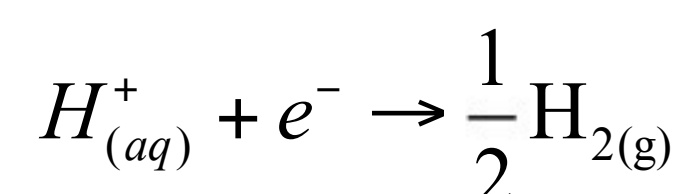
At anode:

NO_3^- Ions will not be able to react with Pt anode as Pt electrodes are inert in nature. Thus, OH^- or NO_3^- ions can be oxidized at the anode. But due to lower discharge potential, OH^- ions will get preference and it will decompose to liberate O_2.
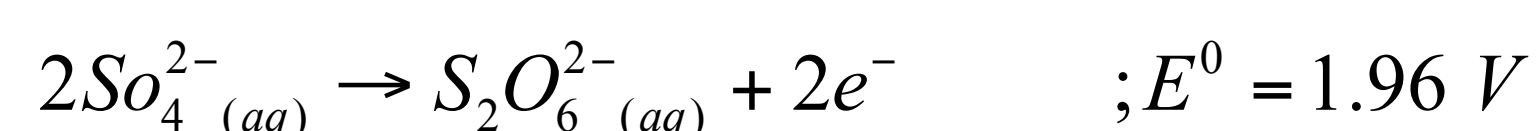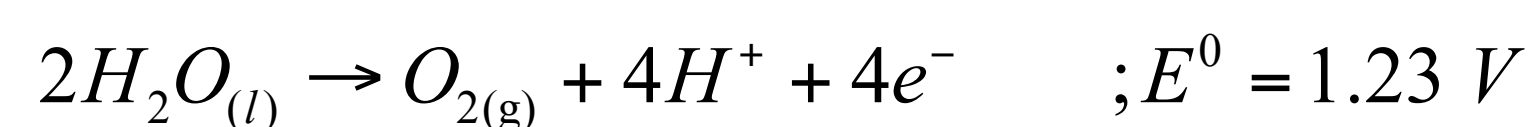
$$OH^- \rightarrow OH + e^-$$

$$4OH^- \rightarrow 2H_2O + O_2$$

iii.

At the cathode: reduction reaction

$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2}H_{2(g)}$$

At the anode: following reactions can take place

$$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+ + 4e^- \qquad ; E^0 = 1.23 \ V$$

$$2So_4^{2-}{}_{(aq)} \rightarrow S_2O_6^{2-}{}_{(aq)} + 2e^- \qquad ; E^0 = 1.96 \ V$$
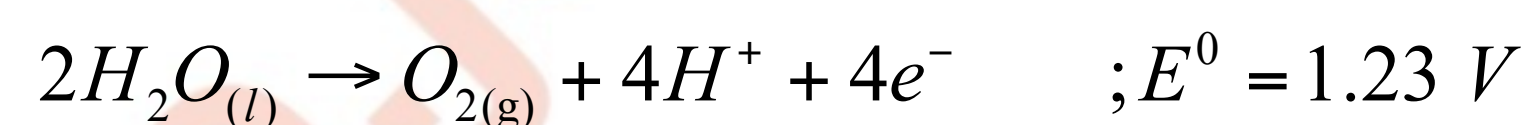
In case of dilute sulphuric acid, first reaction will be preferred & produce O_2 gas. But for concentrated sulphuric acid, second reaction will take place.

iv. At cathode: reduction reactions
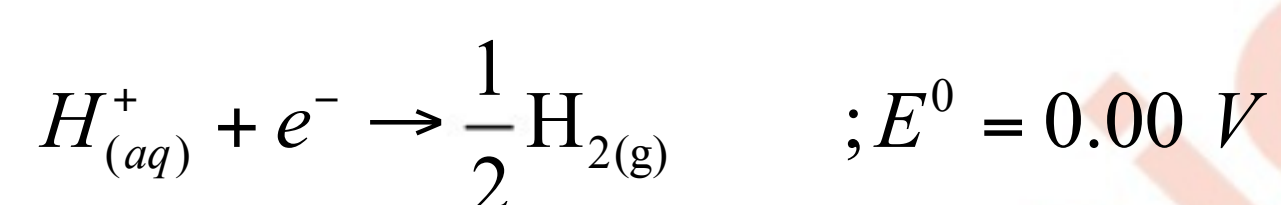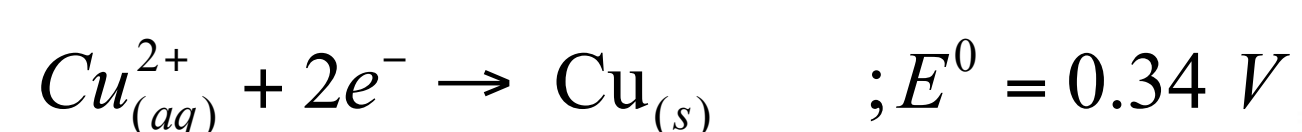
$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad ; E^0 = 0.34 \ V$$

$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2}H_{2(g)} \qquad ; E^0 = 0.00 \ V$$

As we know that the reaction which has higher E^0 value, it will takes place at cathode. Thus, deposition of copper will take place at the cathode.

At anode: following oxidation reactions can take place

$$Cl^-_{(aq)} \rightarrow \frac{1}{2}Cl_{2(g)} + e^- \qquad ; E^0 = 1.36 \ V$$

$$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+ + 4e^- \qquad ; E^0 = 1.23 \ V$$

The reaction which has lower E^0 value, will takes place at anode. Thus, deposition of copper will take place at the cathode. Due to the over-potential of oxygen, Cl^- gets oxidized at the anode to produce Cl_2 gas.