NCERT SOLUTIONS
CLASS-XII CHEMISTRY
CHAPTER-3
ELECTROCHEMISTRY

Q 3.1:

In the order of their reactivity, i.e how they displace each other from their salt solutions, allign the metals in decreasing order. Cu, Fe, Al, Zn and Mg.

Answer:

According to their reactivity, the given metals replace the others from their salt solutions in the said order: Mg, Al, Zn, Fe, Cu .

Mg : Al : Zn : Fe Cu

Q 3.2:

Standard electrode potentials given as,
Mg^{2+}/Mg = −2.37 V, Hg^{2+}/Hg = 0.79V, Cr^{3+}/Cr = − 0.74V, Ag^{+}/Ag = 0.80V, K^{+}/K = −2.93V

In the order of increasing of reducing power arrange the given metals accordingly.

Ans:

The reducing power increases with the lowering of reduction potential. In order of given standard electrode potential (increasing order) : $K^{+}/K < Mg^{2+}/Mg < Cr^{3+}/Cr < Hg^{2+}/Hg < Ag^{+}/Ag$

Thus, in the order of reducing power, we can arrange the given metals as : Ag< Hg < Cr < Mg < K

Q 3.3 :

Represent the galvanic cell in which the following reaction takes place.

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

Also find :

(i) The negatively charged electrode ?

(ii) Current carriers in the cell.

(iii) At each electrode, the individual reaction.

Ans :

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

$$Zn_{(s)}|Zn^{2+}_{(aq)}||Ag^{+}_{(aq)}|Ag_{(s)}$$

(i) The negatively charged electrode is the Zn electrode (anode)

(ii) The current carriers in the cell are ions. Current flows to zinc from silver in the external circuit.

(iii) Reaction at the anode is given by :

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

Reaction at the anode is given by :

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

Q 3.4:

With the following reactions given, find the standard cell potentials of galvanic cells with given reactions.

(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.

Ans :

(i) $E^{\Theta}_{Cr^{3+}/Cr}$ = 0.74 V

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

The galvanic cell of the given reaction is depicted as :

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

Now, the standard cell potential is

$$E^{\Theta}_{cell} = E^{\Theta}_{g} - E^{\Theta}_{L}$$

= – 0.40 – (-0.74)

= + 0.34 V

In the given equation, n = 6

F = 96487 C mol^{-1}

E^{Θ}_{cell} = + 0.34 V

Then, $\Delta_r G^{\Theta}$ = −6 × 96487 C mol^{-1} × 0.34 V

= −196833.48 CV mol−1

= −196833.48 J mol−1

= −196.83 kJ mol−1

Again,

$\Delta_r G^{\Theta} = - RTlnK$ $\Delta_r G^{\Theta} = -2.303RTlnK$ $logk = \frac{\Delta_r G}{2.303RT} = \frac{-196.83\times10^3}{2.303\times8.314\times298}$

= 34.496

K = antilog (34.496) = 3.13 × 10^{34}

The galvanic cell of the given reaction is depicted as:

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

Now, the standard cell potential is

$E^{\Theta}_{cell} = E^{\Theta}_{g} - E^{\Theta}_{L}$

Here, n = 1.

Then, $\Delta_t G^0 = -nFE^0_{cell}$

= −1 × 96487 C mol^{-1} × 0.03 V

= −2894.61 J mol^{-1}

= −2.89 kJ mol^{-1}

Again, $\Delta_t G^0 = -2.303RT\ lnK$

$lnK = \frac{\Delta_t G}{2.303RT} = \frac{-2894.61}{2.303\times8.314\times298}$

= 0.5073

K = antilog (0.5073)

= 3.2 (approximately)

Q 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

(i) For the given reaction, the Nernst equation can be given as:

$E_{cell} = E^0_{cell} - \frac{0.591}{n} log\frac{[Mg^{2+}]}{[Cu^{2+}]} = 0.34 - (-2.36) - \frac{0.0591}{2} log\frac{0.001}{0.0001}$ $2.7 - \frac{0.0591}{2} log10$

= 2.7 − 0.02955

= 2.67 V (approximately)

(ii) For the given reaction, the Nernst equation can be given as:

$E_{cell} = E^0_{cell} - \frac{0.591}{n} log\frac{[Fe^{2+}]}{[H^{+}]^2}$

= 0 – (– 0.14) – $\frac{0.0591}{n} log\frac{0.050}{(0.020)^2}$

= 0.52865 V

= 0.53 V (approximately)

(iii) For the given reaction, the Nernst equation can be given as:

$E_{cell} = E^0_{cell} - \frac{0.591}{n} log \frac{[Sn^{2+}]}{[H^+]^2}$

= 0 – (– 0.14) – $\frac{0.591}{2} log \frac{0.050}{(0.020)^2}$

= 0.14 − 0.0295 × log125

= 0.14 − 0.062

= 0.078 V

= 0.08 V (approximately)

(iv) For the given reaction, the Nernst equation can be given as:

$E_{cell} = E^0_{cell} - \frac{0.591}{n} log \frac{1}{[Br^-]^2[H^+]^2}$

= 0 – 1.09 – $\frac{0.591}{2} log \frac{1}{(0.010)^2(0.030)^2}$

= -1.09 – 0.02955 x $log \frac{1}{0.00000009}$

= -1.09 – 0.02955 x $log \frac{1}{9 \times 10^{-8}}$

= -1.09 – 0.02955 x $log(1.11 \times 10^7)$

= -1.09 – 0.02955 x (0.0453 + 7)

= -1.09 – 0.208

= -1.298 V

Q 3.6:

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

$$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^- \; ; \; E^\ominus = 0.76V$$
$$Ag_2O_{(s)} + H_2O_{(l)} + 2e^- \longrightarrow 2Ag_{(s)} + 2OH^-_{(aq)} \; ; \; E^\ominus = 0.344 \text{ V}$$
$$Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \longrightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)} + 2OH^-_{(aq)} ; \; E^\ominus = 1.104 \text{ V}$$

For the given reaction calculate $\Delta_r G^\Theta$ and E^0 :

Ans:

E^0 = 1.104 V

We know that,

$\Delta_r G^\Theta = -nFE^\Theta$

= −2 × 96487 × 1.04

= −213043.296 J

= −213.04 kJ

Q 3.7:

For the solution of an electrolyte describe its conductivity and molar conductivity. Also put some light on how they vary with concentration.

Answer

Conductivity of a solution is defined as the conductance of a solution of 1 cm in length and area of cross – section 1 sq. cm. Specific conductance is the inverse of resistivity and it is represented by the symbol κ. If ρ is resistivity, then we can write:

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

At any given concentration, the conductivity of a solution is defined as the unit volume of solution kept between two platinum electrodes with the unit area of cross- section at a distance of unit length.

$G = k\frac{a}{l} = k \times 1 = k$ [Since a = 1 , l = 1]

When concentration decreases there will a decrease in Conductivity. It is applicable for both weak and strong electrolyte. 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 –

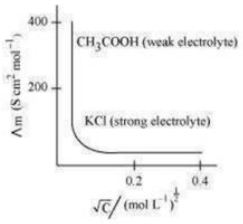
Molar conductivity of a solution at a given concentration is the conductance of volume V of a solution containing 1 mole of the electrolyte kept between two electrodes with the area of cross-section A and distance of unit length.

$\Lambda_m = k\frac{A}{l}$

Now, l = 1 and A = V (volume containing 1 mole of the electrolyte).

$\Lambda_m = kV$

Molar conductivity increases with a decrease in concentration. This is 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 :



Q 3.8:

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

Ans :

Given, κ = 0.0248 S cm^{-1} c

= 0.20 M

Molar conductivity, $\Lambda_m = \frac{k \times 1000}{c}$

$= \frac{0.0248 \times 1000}{0.2}$

= 124 Scm^{2}mol^{-1}

Q 3.9:

Considering the case of a conductivity cell having 0.001 M KCl solution at 298 K is 1500 Ω. If given, conductivity of 0.001M KCl solution at 298 K is 0.146 × 10^{-3} S, find the cell constant?

Answer

Given,

Conductivity, k = 0.146 × 10^{-3} S cm−1

Resistance, R = 1500 Ω

Cell constant = k × R

= 0.146 × 10^{-3} × 1500

= 0.219 cm^{-1}

Q 3.10:

The conductivity of NaCl at 298 K has been found at different concentrations and the results are given below:

Concentration/M 0.001 0.010 0.020 0.050 0.100

10^2 × k/S m^{-1} 1.237 11.85 23.15 55.53 106.74

for all concentrations and draw a plot between Λ_m and c^{½}. Find the value

Molar conductivity of

Calculate Λ_m of Λ_m^0

Ans:

Given,

κ = 1.237 × 10^{-2} S m−1, c = 0.001 M

Then, κ = 1.237 × 10^{-4} S cm^{-1}, c^{½} = 0.0316 M$^{1/2}$

$\Lambda_m = \frac{k}{c} = \frac{1.237 \times 10^{-4} S\ cm^{-1}}{0.001\ mol\ \ L^{-1}} \times \frac{1000\ cm^{-1}}{L}$

= 123.7 S cm^2 mol^{-1}

Given,

$\kappa = 11.85 \times 10^{-2}$ S m^{-1}, c = 0.010M

Then, $\kappa = 11.85 \times 10^{-4}$ S cm^{-1}, $c^{½} = 0.1$ M$^{1/2}$

$\Lambda_m = \frac{k}{c} = \frac{11.85 \times 10^{-4} S\ cm^{-1}}{0.010\ mol\ \ L^{-1}} \times \frac{1000\ cm^{-1}}{L}$

= 118.5 S cm^2 mol^{-1}

Given,

$\kappa = 23.15 \times 10^{-2}$ S m^{-1}, c = 0.020 M

Then, $\kappa = 23.15 \times 10^{-4}$ S cm^{-1}, $c^{1/2} = 0.1414$ M$^{1/2}$

$\Lambda_m = \frac{k}{c} = \frac{23.15 \times 10^{-4} S\ cm^{-1}}{0.020\ mol\ \ L^{-1}} \times \frac{1000\ cm^{-1}}{L}$

= 115.8 S cm^2 mol^{-1}

Given,

$\kappa = 55.53 \times 10^{-2}$ S m^{-1}, c = 0.050 M

Then, $\kappa = 55.53 \times 10^{-4}$ S cm^{-1}, $c^{1/2} = 0.2236$ M$^{1/2}$

$\Lambda_m = \frac{k}{c} = \frac{106.74 \times 10^{-4} S\ cm^{-1}}{0.050\ mol\ \ L^{-1}} \times \frac{1000\ cm^{-1}}{L}$

= 111.1 1 S cm^2 mol^{-1}

Given,

$\kappa = 106.74 \times 10^{-2}$ S m^{-1}, c = 0.100 M

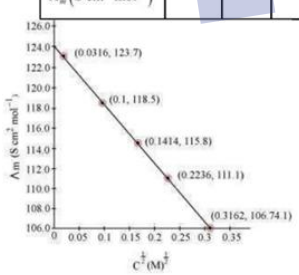
Then, $\kappa = 106.74 \times 10^{-4}$ S cm^{-1}, $c^{1/2} = 0.3162$ M$^{1/2}$

$\Lambda_m = \frac{k}{c} = \frac{106.74 \times 10^{-4} S\ cm^{-1}}{0.100\ mol\ \ L^{-1}} \times \frac{1000\ cm^{-1}}{L}$

= 106.74 S cm^2 mol^{-1}

Now, we have the following data :

$C^{1/2}/M^{1/2}$	0.0316	0.1	0.1414	0.2236	0.3162
Λ_m (S cm^2 mol^{-1})	123.7	118.5	115.8	111.1	106.74



Since the line interrupts Λ_m at 124.0 S cm^2 mol^{-1}, Λ_m^0 = 124.0 S cm^2 mol^{-1}

Q 3.11:

Find the molar conductivity of acetic acid if its conductivity is given to be 0.00241 M . Also, if the value of Λ_m^0 is given to be390.5 S cm^2 mol^{-1}, calculate its dissociation constant?

Ans:

Given, $\kappa = 7.896 \times 10^{-5}$ S m^{-1} c

= 0.00241 mol L^{-1}

Then, molar conductivity, $\Lambda_m = \frac{k}{c}$

= $\frac{7.896 \times 10^{-5} Scm^{-1}}{0.00241\ mol\ L^{-1}} \times \frac{1000cm^3}{L}$

= 32.76S cm^2 mol^{-1}

$\Lambda_m^0 =$ 390.5 S cm^2 mol^{-1}

Again,

$\alpha = \frac{\Lambda_m}{\Lambda_m^0}$

$= = \frac{32.76\ S\ cm^2\ mol^{-1}}{390.5\ S\ cm^2\ mol^{-1}}$

Now,

= 0.084

Dissociation constant, $K_a = \frac{c\alpha^2}{(1-\alpha)}$

$= \frac{(0.00241\ mol\ L^{-1})(0.084)^2}{(1-0.084)}$

= 1.86×10^{-5} mol L^{-1}

Q 3.12:

How much charge is required for the following reductions of 1 mol of :

(i) Al^{3+} to Al.

(ii) Cu^{2+} to Cu.

(iii)MnO_4^- to Mn^{2+}.

Ans :

(i) $Al^{3+} + 3e^- \rightarrow Al$

Required charge = 3 F

= 3 × 96487 C

= 289461 C

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

Required charge = 2 F

= 2 × 96487 C

= 192974 C

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

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

Required charge = 5 F

= 5 × 96487 C

= 482435 C

Q 3.13:

In the terms of Faraday, how much electricity is required to produce :

(i) From molten $CaCl_2$, 20.0 g of Ca.

(ii) From molten Al_2O_3, 40.0 g of Al.

Ans:

(i) From given data,

$Ca^{2+} + 2e^- \rightarrow Ca$

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium = (2 x 20)/ 40 F

= 1 F

(ii) From given data,

$Al^{3+} + 3e^- \rightarrow Al$

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = (3 x 40)/27 F

= 4.44 F

Q 3.14:

Calculate the amount of electricity required for the oxidation of 1 mol of the following in coulombs :

(i) H_2O to O_2.

(ii)FeO to Fe_2O_3.

Ans :

(i) From given data,

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

We can say that :

$$O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$$

Electricity required for the oxidation of 1 mol of H_2O to O_2 = 2 F

= 2 × 96487 C

= 192974 C

(ii) From given data,

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

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

= 96487 C

Q 3.15:

For 20 minutes, a current of 5 A is applied to between platinum electrodes to electrolyze a solution of $Ni(NO_3)_2$. Find the amount of Ni deposited at the cathode?

Ans :

Given,

Current = 5A

Time = 20 × 60 = 1200 s

Charge = current × time

= 5 × 1200

= 6000 C

According to the reaction,

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

Nickel deposited by 2 × 96487 C = 58.71 g

Therefore, nickel deposited by 6000 C = $\frac{58.71 \times 6000}{2 \times 96487} g$

= 1.825 g

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

Q 3.16:

Solutions of 3 electrolytic cells are $ZnSO_4$, $AgNO_3$ and $CuSO_4$,cells are connected in series. Of the cells, A,B,C respectively, after passing a steady current of 1.5 amperes , 1.45 g of silver was found deposited at the cathode of cell B. How much time did the current flow? What amount of zinc and copper were deposited?

Ans :

According to the reaction:

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

i.e., 108 g of Ag is deposited by 96487 C.

Therefore, 1.45 g of Ag is deposited by = $\frac{96487 \times 1.45}{107} C$

= 1295.43 C

Given,

Current = 1.5 A

Time = 1295.43/ 1.5 s

= 863.6 s

= 864 s

= 14.40 min

Again,

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

i.e., 2 × 96487 C of charge deposit = 63.5 g of Cu

Therefore, 1295.43 C of charge will deposit $\frac{63.5 \times 1295.43}{2 \times 96487}$

= 0.426 g of Cu

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

i.e., 2 × 96487 C of charge deposit = 65.4 g of Zn

Therefore, 1295.43 C of charge will deposit $\frac{65.4 \times 1295.43}{2 \times 96487}$

= 0.439 g of Zn

Q 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) and I^-(aq)

(ii) Ag^+ (aq) and Cu(s)

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

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

(v) Br_2 (aq) and Fe^{2+} (aq).

Ans :

(i)

$$\begin{array}{lll} Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} \;] \times 2; & & E^\circ = +0.77\ V \\ 2I^-_{(aq)} \longrightarrow I_{2(s)} + 2e^-; & & E^\circ = -0.54\ V \\ \hline 2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)} & ; & E^\circ = +0.23\ V \end{array}$$

E° is positive , hence reaction is feasible.

(ii)

$$\begin{array}{lll} Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)} \;] \times 2 & ; & E^\circ = +0.80\ V \text{ Fe} \\ Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^- & ; & E^\circ = -0.34\ V \\ \hline 2Ag^+_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)} & ; & E^\circ = +0.46\ V \end{array}$$

E^0 is positive , hence reaction is feasible.

(iii)

$$\begin{array}{lll} Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} \;] \times 2 & ; & E^\circ = +0.77\ V \\ 2Br^-_{(aq)} \longrightarrow Br_{2(l)} + 2e^- & ; & E^\circ = -1.09\ V \\ \hline 2Fe^{3+}_{(aq)} + 2Br^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} \text{ and } Br_{2(l)} & ; & E^\circ = -0.32\ V \end{array}$$

E^0 is negative , hence reaction is not feasible.

(iv)

$$\begin{array}{lll} Ag_{(s)} \longrightarrow Ag^+_{(aq)} + e^- & ; & E^\circ = -0.80\ V \\ Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} & ; & E^\circ = +0.77\ V \\ \hline Ag + Fe^{3+} \rightarrow Ag^+ + Fe^{2+} & ; & E^\circ = -0.03\ V \end{array}$$

E^0 is negative , hence reaction is not feasible.

(v)

$$Br_{2(aq)} + 2e^- \longrightarrow 2Br^-_{(aq)} \quad ; \quad E^\circ = +1.09\ V$$
$$Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^- \ ] \times 2 \quad ; \quad E^\circ = -0.77\ V$$
$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)} \quad ; \quad E^\circ = +0.32\ V$$

E^0 is positive , hence reaction is feasible.

Q 3.18:

Predict the products of electrolysis in each of the following :

(i) An aqueous solution of $AgNO_3$ with silver electrodes.

(ii) An aqueous solution of $AgNO_3$with platinum electrodes.

(iii) A dilute solution of $H_2SO_4$with platinum electrodes.

(iv) An aqueous solution of $CuCl_2$ with platinum electrodes.

Ans:

(i) At cathode:

The following reduction reactions compete to take place at the cathode.

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

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

The reaction with a higher value of E^0 takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

The Ag anode is attacked by NO_3^+ ions. Therefore, the silver electrode at the anode dissolves in the solution to form Ag^+.

(ii) At cathode:

The following reduction reactions compete to take place at the cathode.

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

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

The reaction with a higher value of E^0 takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

Since Pt electrodes are inert, the anode is not attacked by NO_3^+ ions. Therefore, OH^- or NO_3^+ ions can be oxidized at the anode. But OH^- ions having a lower discharge potential and get preference and decompose to liberate O_2.

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

(iii) At the cathode, the following reduction reaction occurs to produce H_2 gas.

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

At the anode, the following processes are possible.

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

$2SO^{2-}_{4(aq)} \rightarrow S_2O^{2-}_{6(aq)} + 2e^-$; E^0 = +1.96 V —(ii)

For dilute sulphuric acid, reaction (i) is preferred to produce O_2 gas. But for concentrated sulphuric acid, reaction (ii) occurs.

(iv) At cathode:

The following reduction reactions compete to take place at the cathode.

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

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

The reaction with a higher value of takes place at the cathode. Therefore, deposition of copper will take place at the cathode.

At anode:

The following oxidation reactions are possible at the anode.

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

$2H_20_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + e^-$; E^0 = +1.23 V

At the anode, the reaction with a lower value of E^0 is preferred. But due to the over potential of oxygen, Cl^- gets oxidized at the anode to produce Cl_2 gas.