

Q 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:

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:

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.

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 :

Depict the galvanic cell in which the reaction

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

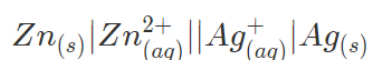
(i) Which of the electrode is negatively charged?

(ii) The carriers of the current in the cell.

(iii) Individual reaction at each electrode.

Ans :

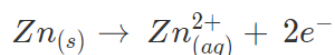
The galvanic cell in which the given reaction takes place is depicted as:



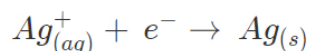
(i) The negatively charged electrode is the Zn electrode (anode)

(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 :



Reaction at the anode is given by :



Q 3.4:

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

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

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

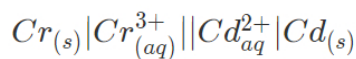
Calculate the $\Delta_r G^\circ$ and equilibrium constant of the reactions.

Ans :

$$(i) E_{Cr^{3+}/Cr}^{\ominus} = 0.74 \text{ V}$$

$$E_{Cd^{2+}/Cd}^{\ominus} = -0.40 \text{ V}$$

The galvanic cell of the given reaction is depicted as :



Now, the standard cell potential is

$$E_{cell}^{\ominus} = E_g^{\ominus} - E_L^{\ominus}$$

$$= -0.40 - (-0.74)$$

$$= +0.34 \text{ V}$$

In the given equation, n = 6

$$F = 96487 \text{ C mol}^{-1}$$

$$E_{cell}^{\ominus} = +0.34 \text{ V}$$

$$\text{Then, } \Delta_r G^{\ominus} = -6 \times 96487 \text{ C mol}^{-1} \times 0.34 \text{ V}$$

$$= -196833.48 \text{ CV mol}^{-1}$$

$$= -196833.48 \text{ J mol}^{-1}$$

$$= -196.83 \text{ kJ mol}^{-1}$$

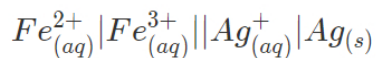
Again,

$$\Delta_r G^{\ominus} = -RT \ln K \quad \Delta_r G^{\ominus} = -2.303RT \log k = \frac{\Delta_r G}{2.303RT} = \frac{-196.83 \times 10^3}{2.303 \times 8.314 \times 298}$$

$$= 34.496$$

$$K = \text{antilog}(34.496) = 3.13 \times 10^{34}$$

The galvanic cell of the given reaction is depicted as:



Now, the standard cell potential is

$$E_{cell}^{\ominus} = E_g^{\ominus} - E_L^{\ominus}$$

Here, n = 1.

$$\text{Then, } \Delta_t G^0 = -nFE_{cell}^0$$

$$= -1 \times 96487 \text{ C mol}^{-1} \times 0.03 \text{ V}$$

$$= -2894.61 \text{ J mol}^{-1}$$

$$= -2.89 \text{ kJ mol}^{-1}$$

$$\text{Again, } \Delta_t G^0 = -2.303RT \ln K \quad \ln K = \frac{\Delta_t G}{2.303RT} = \frac{-2894.61}{2.303 \times 8.314 \times 298}$$

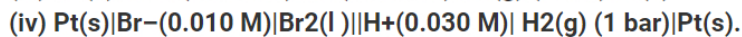
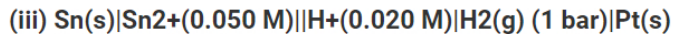
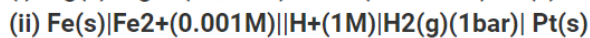
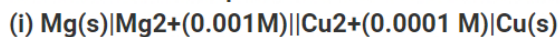
$$= 0.5073$$

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

$$= 3.2 \text{ (approximately)}$$

Q 3.5:

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



Answer

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

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.591}{n} \log \frac{[\text{Mg}^{2+}]}{[\text{Cu}^{2+}]} = 0.34 - (-2.36) - \frac{0.0591}{2} \log \frac{0.001}{0.0001} = 2.7 - \frac{0.0591}{2} \log 10$$

$$= 2.7 - 0.02955$$

$$= 2.67 \text{ V (approximately)}$$

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

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

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

$$= 0.52865 \text{ V}$$

$$= 0.53 \text{ V (approximately)}$$

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

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

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

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

$$= 0.14 - 0.062$$

$$= 0.078 \text{ V}$$

$$= 0.08 \text{ V (approximately)}$$

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

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

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

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

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

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

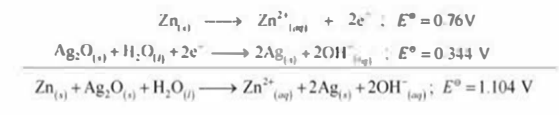
$$= -1.09 - 0.02955 \times (0.0453 + 7)$$

$$= -1.09 - 0.208$$

$$= -1.298 \text{ V}$$

Q 3.6:

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



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

Ans:

$$E^\ominus = 1.104 \text{ V}$$

We know that,

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

$$= -2 \times 96487 \times 1.04$$

$$= -213043.296 \text{ J}$$

$$= -213.04 \text{ kJ}$$

Q 3.7:

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

Answer

The 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:

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

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.

$$G = \kappa \frac{a}{l} = \kappa \times 1 = \kappa \quad [\text{Since } a = 1, l = 1]$$

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.

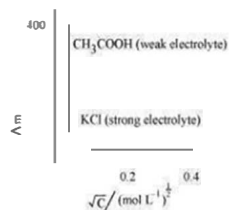
$$\Lambda_m = \kappa \frac{A}{l}$$

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 S cm⁻¹. Calculate its molar conductivity

Ans :

$$\text{Given, } \kappa = 0.0248 \text{ S cm}^{-1} \text{ c}$$

$$= 0.20 \text{ M}$$

$$\text{Molar conductivity, } \Lambda_m = \frac{k \times 1000}{c} = \frac{0.0248 \times 1000}{0.2}$$

$$= 124 \text{ Scm}^2\text{mol}^{-1}$$

Q 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

Given,

$$\text{Conductivity, } k = 0.146 \times 10^{-3} \text{ S cm}^{-1}$$

$$\text{Resistance, } R = 1500 \Omega$$

$$\text{Cell constant} = k \times R$$

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

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

Q 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
$10^2 \times k/\text{S m}^{-1}$	1.237	11.85	23.15	55.53	106.74

Calculate Λ_m for all concentrations and draw a plot between Λ_m and $c^{1/2}$. Find the value of Λ_m° .

Ans:

Given,

$$\kappa = 1.237 \times 10^{-2} \text{ S m}^{-1}, c = 0.001 \text{ M}$$

$$\text{Then, } \kappa = 1.237 \times 10^{-4} \text{ S cm}^{-1}, c^{1/2} = 0.0316 \text{ M}^{1/2}$$

$$\Lambda_m = \frac{k}{c} = \frac{1.237 \times 10^{-4} \text{ S cm}^{-1}}{0.001 \text{ mol L}^{-1}} \times \frac{1000 \text{ cm}^3}{\text{L}}$$

$$= 123.7 \text{ S cm}^2 \text{ mol}^{-1}$$

Given,

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

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

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

$$= 118.5 \text{ S cm}^2 \text{ mol}^{-1}$$

Given,

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

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

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

$$= 115.8 \text{ S cm}^2 \text{ mol}^{-1}$$

Given,

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

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

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

$$= 111.1 \text{ S cm}^2 \text{ mol}^{-1}$$

Given,

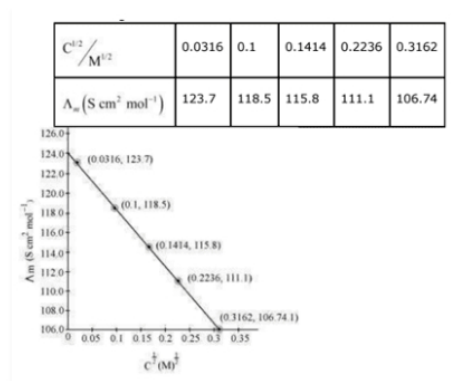
$$\kappa = 106.74 \times 10^{-2} \text{ S m}^{-1}, c = 0.100 \text{ M}$$

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

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

$$= 106.74 \text{ S cm}^2 \text{ mol}^{-1}$$

Now, we have the following data :



Since the line intercepts Λ_m at $124.0 \text{ S cm}^2 \text{ mol}^{-1}$, $\Lambda_m^0 = 124.0 \text{ S cm}^2 \text{ mol}^{-1}$

Q 3.11:

Conductivity of 0.00241 M acetic acid is $7.896 \times 10^{-5} \text{ S cm}^{-1}$. Calculate its molar conductivity. If $0 \Lambda_m$ for acetic acid is $390.5 \text{ S cm}^2 \text{ mol}^{-1}$, what is its dissociation constant?

Ans:

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

$$= 0.00241 \text{ mol L}^{-1}$$

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

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

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

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

Again,

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

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

Now,

$$= 0.084$$

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

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

$$= 1.86 \times 10^{-5} \text{ mol L}^{-1}$$

Q 3.12:

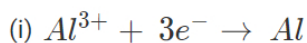
How much charge is required for the following reductions:

(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+} ?

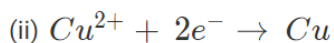
Ans :



Required charge = 3 F

$$= 3 \times 96487 \text{ C}$$

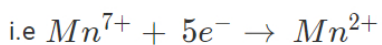
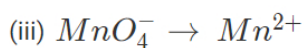
$$= 289461 \text{ C}$$



Required charge = 2 F

$$= 2 \times 96487 \text{ C}$$

$$= 192974 \text{ C}$$



Required charge = 5 F

$$= 5 \times 96487 \text{ C}$$

$$= 482435 \text{ C}$$

Q 3.13:

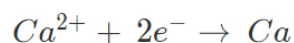
How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten CaCl_2 ?

(ii) 40.0 g of Al from molten Al_2O_3 ?

Ans:

(i) From given data,

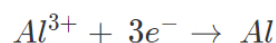


Electricity required to produce 40 g of calcium = 2 F

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

$$= 1 \text{ F}$$

(ii) From given data,



Electricity required to produce 27 g of Al = 3 F

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

$$= 4.44 \text{ F}$$

Q 3.14:

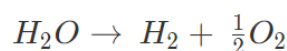
How much electricity is required in coulomb for the oxidation of

(i) 1 mol of H_2O to O_2 ?

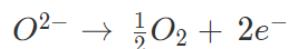
(ii) 1 mol of FeO to Fe_2O_3 ?

Ans :

(i) From given data,



We can say that :



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

$$= 2 \times 96487 \text{ C}$$

$$= 192974 \text{ C}$$

(ii) From given data,



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

$$= 96487 \text{ C}$$

Q 3.15:

A solution of $\text{Ni}(\text{NO}_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Ans :

Given,

Current = 5A

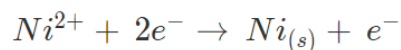
Time = 20 × 60 = 1200 s

Charge = current × time

= 5 × 1200

= 6000 C

According to the reaction,



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

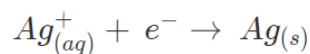
Hence, 1.825 g of nickel will be deposited at the cathode.

Q 3.16:

Three electrolytic cells A,B,C containing solutions of ZnSO₄, AgNO₃ and CuSO₄, 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?

Ans :

According to the reaction:



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

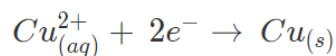
Time = 1295.43/ 1.5 s

= 863.6 s

= 864 s

= 14.40 min

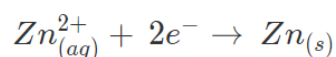
Again,



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



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

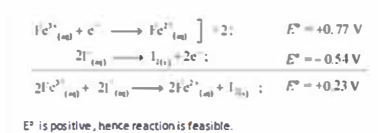
Q 3.17:

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

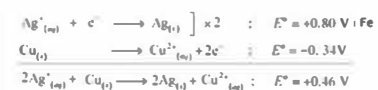
- (i) $\text{Fe}^{3+}(\text{aq})$ and $\text{I}^{-}(\text{aq})$
- (ii) $\text{Ag}^{+}(\text{aq})$ and $\text{Cu}(\text{s})$
- (iii) $\text{Fe}^{3+}(\text{aq})$ and $\text{Br}^{-}(\text{aq})$
- (iv) $\text{Ag}(\text{s})$ and $\text{Fe}^{3+}(\text{aq})$
- (v) $\text{Br}_2(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$.

Ans :

(i)

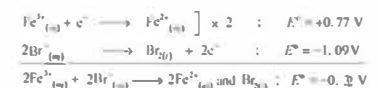


(ii)



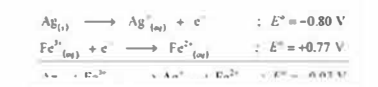
E° is positive, hence reaction is feasible.

(iii)



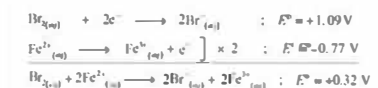
E° is negative, hence reaction is not feasible.

(iv)



E° is negative, hence reaction is not feasible.

(v)



E° is positive, hence reaction is feasible.

Q 3.18:

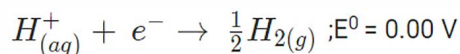
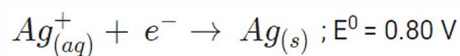
Predict the products of electrolysis in each of the following:

- (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.



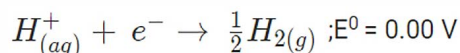
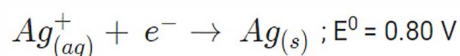
The reaction with a higher value of E^0 takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

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.



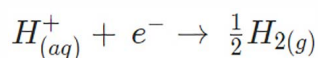
The reaction with a higher value of E^0 takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

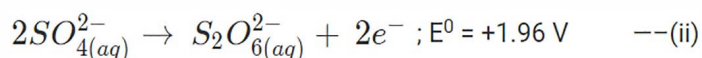
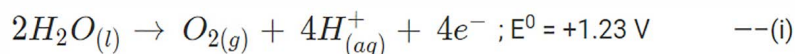
Since Pt electrodes are inert, the anode is not attacked by NO_3^{+} ions. Therefore, OH^{-} or NO_3^{+} ions can be oxidized at the anode. But OH^{-} ions having a lower discharge potential and get preference and decompose to liberate O_2 .



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



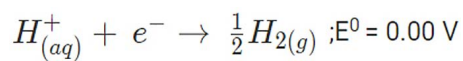
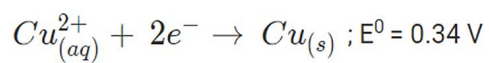
At the anode, the following processes are possible.



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

(iv) At cathode:

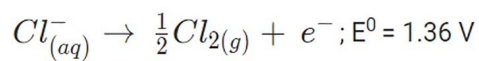
The following reduction reactions compete to take place at the cathode.



The reaction with a higher value of E^0 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.



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.