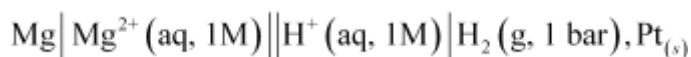


Question 3.1 How would you determine the standard electrode potential of the system $Mg^{2+}|Mg$?

Answer :

To determine the standard electrode potential of the given system we need to use a hydrogen electrode. In the setup, we shall put a hydrogen electrode as cathode and $Mg|MgSO_4$ as an anode.



Now we will measure the emf of the cell. This emf will be the standard electrode potential of the magnesium electrode.

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{right}} - E^{\circ}_{\text{left}}$$

$$E^{\circ}_{\text{left}} = 0 \text{ (The standard hydrogen electrode is always zero)}$$

Hence

$$E^{\circ}_{\text{cell}} = E^{\circ}_{Mg|Mg^{2+}}$$

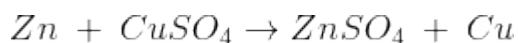
Question 3.2 Can you store copper sulphate solutions in a zinc pot?

Answer:

The standard electrode potential of Zinc is - 0.76 whereas that of Copper is 0.34. So Zinc will reduce copper into the lower state.

It is known that zinc is more reactive than copper. Thus if we will store copper sulphate solution in zinc pot then zinc will displace copper from its solution.

The following reaction will take place:-



Question 3.3 Consult the table of standard electrode potentials and suggest three substances that can oxidise ferrous ions under suitable conditions.

Answer :

The oxidising strength of elements increases as the standard electrode potential increases.

So all the elements having greater standard potential than iron can oxidise it to a higher state.

Few such elements are :- F_2 , Cl_2 , Br_2 , Ag^{+1} etc.

Question 3.4 Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

Answer :

It is given that pH of the solution is 10, i.e., the hydrogen ion concentration in the solution is 10^{-10} M.



By Nernst equation we have :-

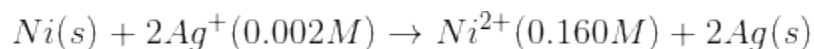
$$E_{Cell} = E_{cell}^{\circ} - \frac{RT}{2F} \ln \frac{1}{[H^+]}$$

$$\text{So, } = 0 - \frac{0.0591}{1} \log \frac{1}{[10^{-10}]}$$

$$\text{or } = - 0.591 \text{ V}$$

So the required potential is - 0.591 V.

Question 3.5 Calculate the emf of the cell in which the following reaction takes place:



Given that $E_{(cell)}^{\ominus} = 1.05 V$

Answer :

Here we can directly apply the nernst equation :-

$$E_{Cell} = E_{cell}^{\circ} - \frac{0.0591}{n} \log \frac{[Ni^{+2}]}{[Ag^+]^2}$$

Putting the value in this equation :-

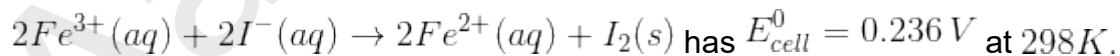
$$= 1.05 - \frac{0.0591}{2} \log \frac{0.160}{(0.002)^2}$$

$$\text{or} = 1.05 - 0.02955 \log(4 \times 10^4)$$

$$\text{or} = 0.914 V$$

Hence the required potential is 0.914 V.

Question 3.6 The cell in which the following reaction occurs:



Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

Answer :

For finding Gibbs free energy we know the relation :-

$$\Delta G_r^\circ = - nFE_{cell}^\circ$$

$$= - 2 \times 96487 \times 0.236$$

$$= - 45541.864 \text{ J mol}^{-1}$$

$$= - 45.54 \text{ KJ mol}^{-1}$$

Now, for equilibrium constant we will use :-

$$\Delta G_r^\circ = -2.303 RT \log K_c$$

$$\text{So, } \log K_c = - \frac{-45.54 \times 10^3}{2.303 \times 8.314 \times 298}$$

$$\text{or } \log K_c = 7.981$$

$$\text{or } K_c = 9.57 \times 10^7$$

Question 3.7 Why does the conductivity of a solution decrease with dilution?

Answer :

The conductivity of a solution depends upon the number of ions and the distance between them. In the process of dilution, we don't increase the number of ions in the solution instead we increase the distance between them. So the conductivity of the solution decreases due to dilution.

Question 3.8 Suggest a way to determine the Λ_m^0 value of water.

Answer :

We know :

$$\Lambda_m = \Lambda_m^\circ - Ac^{\frac{1}{2}}$$

If we draw a straight line between Λ_m and $\sqrt{c}$, its slope will be -A and the intercept on the y-axis will be Λ_m° .

In this way, we can obtain the value of limiting molar conductivity.

Question 3.8 The molar conductivity of 0.025 mol L^{-1} methanoic acid is $46.0 \text{ S cm}^2 \text{ mol}^{-1}$

Calculate its degree of dissociation and dissociation constant.

Given $\lambda^\circ(H^+) = 349.6 \text{ S cm}^2 \text{ mol}^{-1}$ and $\lambda^\circ(HCOO^-) = 54.6 \text{ S cm}^2 \text{ mol}^{-1}$

Answer :

We know that :-

$$\Lambda_m = \lambda^\circ(H^+) + \lambda^\circ(HCOO^-)$$

$$= 349.6 + 54.6$$

$$= 404.2 \text{ S cm}^2 \text{ mol}^{-1}$$

For degree of dissociation, we have :-

$$\alpha = \frac{\Lambda_m(HCOOH)}{\Lambda^\circ(HCOOH)}$$

$$\text{or } \alpha = \frac{46.1}{404.2} = 0.114$$

For dissociation constant, we have :-

$$K_a = \frac{c\alpha^2}{1 - \alpha}$$

$$\text{or } K_a = \frac{0.025 \times (0.114)^2}{1 - 0.114}$$

$$\text{or } = 3.67 \times 10^{-4} \text{ mol L}^{-1}$$

Question 3.10 If a current of 0.5 ampere flows through a metallic wire for 2 hours, then how many electrons would flow through the wire?

Answer :

Firstly we will find total charge flown through the wire then we will calculate number of electrons.

We are given :- $I = 0.5 \text{ A}$, Time = 2 hours = 7200 seconds.

We have, $Q = I.t$

$$= (0.5)7200 = 3600 \text{ C.}$$

Now we will convert charge into number of electrons.

We know that $96487 \text{ C} = 6.023 \times 10^{23} \text{ No. of electrons}$

So total number of electrons :

$$= \frac{3600}{96487} \times 6.023 \times 10^{23}$$

$$\text{or } = 2.25 \times 10^{22} \text{ no. of electrons will flow through wire.}$$

Question 3.11 Suggest a list of metals that are extracted electrolytically.

Answer :

Metals like Na, Mg, Al, etc. are produced on a large scale by electrochemical reduction of their respective cations or by the process electrolysis because there are no suitable reducing agents available for this purpose.

Question 3.12 Consider the reaction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

What is the quantity of electricity in coulombs needed to reduce 1 mol of $Cr_2O_7^{2-}$?

Answer :

It is clear from the given reaction that reduction of 1 mol of $Cr_2O_7^{2-}$ will be

= 6 F (as 6 electrons are required to balance the reaction; Charge required = nF)

= 6×96500

= 579000 C

Thus 578922 C charge is required for reduction of 1 mol of $Cr_2O_7^{2-}$.

Question 3.13 Write the chemistry of recharging the lead storage battery, highlighting all the materials that are involved during recharging.

Answer :

The lead storage battery can be recharged by reversing the direction of current passing through it.

For recharging $PbSO_4$ is converted into Pb at the anode and into PbO_2 at the cathode.

The chemical reactions are as follows:-



Question 3.14 Suggest two materials other than hydrogen that can be used as fuels in fuel cells.

Answer :

The two materials are methane and methanol that can be used as fuels in fuel cells.

Question 3.15 Explain how rusting of iron is envisaged as setting up of a electrochemical cell.

Answer :

The chemistry of corrosion is quite complex but it can be understood by considering it as an electrochemical phenomenon. Consider a particular spot on an object where corrosion takes place. At here oxidation takes place and this spot behaves as the anode. The released electrons at anodic spot go through the metal and go to another spot on the metal and reduction of oxygen takes place in the presence of H^+ . This spot behaves as a cathode with the reaction. In this way this analogy is possible.

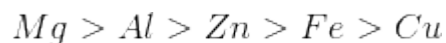
Question 3.1 Arrange the following metals in the order in which they displace each other from the solution of their salts.

Al, Cu, Fe, Mg and Zn.

Answer :

The order in which metals displace each other from the solution of their salts can be given with the help of their standard electrode potential. Since magnesium has the least

standard electrode potential so it is the most strong reducing agent. So the required order we get is:-



Question 3.2 Given the standard electrode potentials,

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

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

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

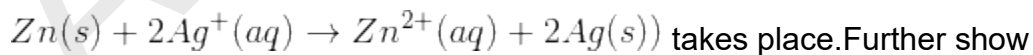
Arrange these metals in their increasing order of reducing power.

Answer :

Elements with reducing power or reducing agents have least/minimum standard electrode potential i.e., reducing power increases with a decrease in standard electrode potential. So the result obtained is:-



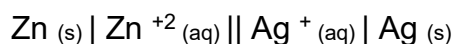
Question 3.3(i) Depict the galvanic cell in which the reaction



(i) Which of the electrode is negatively charged?

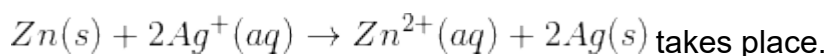
Answer :

The galvanic cell of the given reaction is depicted below :-



Clearly Zn electrode is negatively charged.

Question 3.3(ii) Depict the galvanic cell in which the reaction

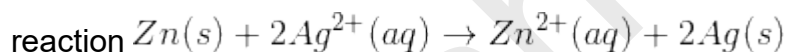


(ii) Further show: The carriers of the current in the cell.

Answer :

The carriers of current in the cell are **ions** . and Current flows from silver to zinc in the external circuit.

Question 3.3(iii) Depict the galvanic cell in which the



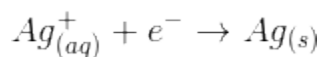
takes place.

(iii) Further show: Individual reaction at each electrode.

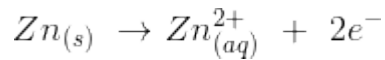
Answer :

The reaction taking place at both cathode and anode are shown below :-

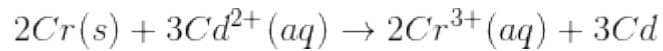
(i) Cathode reaction :-



(ii) Anode reaction :-



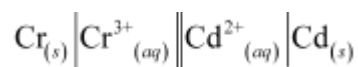
Question 3.4(i) Calculate the standard cell potentials of galvanic cell in which the following reactions take place:



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

Answer :

The galvanic cell of the given reaction is shown below:-



The standard electrode potential of Cr and Cd can be found in table of standard electrode potential.

So, we get :

$$\begin{aligned} E^\circ &= E^\circ_R - E^\circ_L \\ &= -0.40 - (-0.74) \\ &= 0.34 \text{ V} \end{aligned}$$

Now

$$\Delta_r G^\circ = - n F E^\circ_{\text{cell}}$$

Putting values :

$$\Delta_r G^\circ = - 6 \times 96487 \times 0.34$$

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

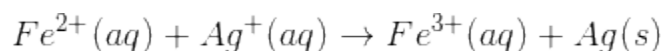
Now for finding equilibrium constant we have :

$$\log k = \frac{-\Delta G_r^\circ}{2.303 \times R \times T}$$

$$\text{or } \log k = 34.496$$

$$\text{or } K = 3.13 \times 10^{34}$$

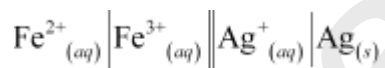
Question 3.4(ii) Calculate the standard cell potentials of galvanic cell in which the following reactions take place:



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

Answer :

The galvanic cell of the given reaction is shown below :-



We can know about the electrode potential of Fe and Ag with the help of table of standard electrode potential.

$$\text{We have : } E^\circ_{(cell)} = E^\circ_R - E^\circ_L$$

$$\text{or } = 0.80 - 0.77$$

$$\text{or } = 0.03 \text{ V}$$

$$\text{Now consider : } \Delta G_r^\circ = - n F E^\circ_{(cell)}$$

$$\text{or} = -1 \times 96487 \times 0.03$$

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

Now for equilibrium constant :

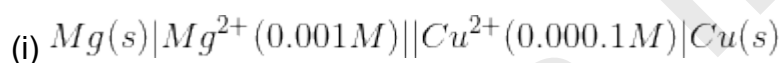
$$\log K = -\frac{\Delta G_r^\circ}{2.303 \times RT}$$

$$\text{or} = -\frac{-2894.61}{2.303 \times 8.314 \times 298}$$

$$\text{or} = 0.5073$$

$$\text{Thus } k \approx 3.2$$

Question 3.5(i) Write the Nernst equation and emf of the following cells at 298 K:



Answer :

The nernst equation gives :

$$E_{Cell} = E_{cell}^\circ - \frac{0.059}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$$

This gives,

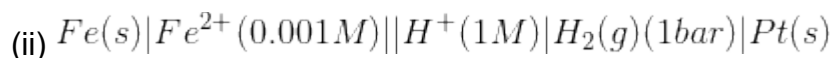
$$= 0.34 - (-2.36) - \frac{0.059}{2} \log \frac{0.001}{0.0001}$$

$$\text{or} = 2.7 - 0.02955$$

$$\text{or} = 2.67 \text{ V}$$

So the emf of the cell is 2.67 V.

Question 3.5(ii) Write the Nernst equation and emf of the following cells at 298 K:



Answer :

The nernst equation for this gives :

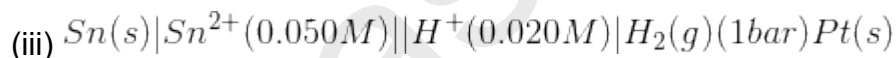
$$E_{Cell} = E_{cell}^{\circ} - \frac{0.0591}{n} \log \frac{[Fe^{+2}]}{[H^+]^2}$$

This gives : $= 0 - (-0.44) - \frac{0.0591}{2} \log \frac{0.001}{1^2}$

or $= 0.44 - 0.02955(-3) = 0.53 \text{ V}$

Thus the emf of the given galvanic cell is 0.53 V.

Question 3.5(iii) Write the Nernst equation and emf of the following cells at 298 K:



Answer :

The nernst equation for this reaction gives :-

$$E_{Cell} = E_{cell}^{\circ} - \frac{0.0591}{n} \log \frac{[Sn^{+2}]}{[H^+]^2}$$

Now for emf, just put all the values.

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

$$\text{or} = 0.14 - 0.0295 \times \log 125$$

$$\text{or} = 0.14 - 0.062 = 0.078 \text{ V}$$

Thus emf of the cell is 0.078 V.

Question 3.5(iv) Write the Nernst equation and emf of the following cells at 298 K:



Answer :

The nernst equation of the given reaction gives :

$$E_{\text{Cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{1}{[\text{Br}^-]^2 [\text{H}^+]^2}$$

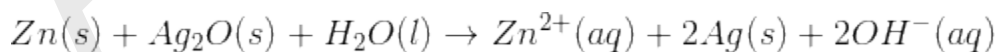
$$\text{or} = (0 - 1.09) - \frac{0.0591}{2} \log \frac{1}{(0.010)^2 (0.030)^2}$$

$$\text{or} = -1.09 - 0.02955 \times \log(1.11 \times 10^7)$$

$$\text{or} = -1.09 - 0.208 = -1.298 \text{ V}$$

So the required emf of the cell is -1.298 V.

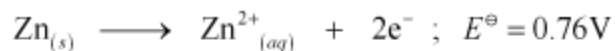
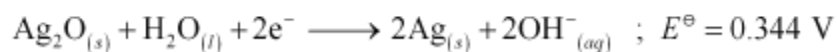
Question 3.6 In the button cells widely used in watches and other devices the following reaction takes place:



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

Answer :

The given reaction is obtained from :-



So the E°_{cell} can be obtained directly.

$$E^{\circ}_{\text{cell}} = 0.76 - (-0.344) = 1.104\text{ V}$$

Now for free energy calculation, we have :-

$$\Delta G^{\circ}_r = -nFE^{\circ}_{\text{cell}}$$

$$\text{or} = -2 \times 96487 \times 1.04$$

$$\text{or} = -213043.29\text{ J}$$

$$\text{or} = -213.04\text{ KJ}$$

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

Answer :

Conductivity(k) or specific conductance of a solution is defined as the inverse of resistivity.

Mathematically, it can be written as:-

$$G = \kappa \frac{A}{L}$$

In the above equation is κ the conductivity of a solution. Thus the definition of conductivity becomes as the conductance of a substance which is 1 cm long and has 1 sq. cm of cross-sectional area.

With dilution conductivity of a solution **decreases** due to an increase in distance between ions.\

Molar conductivity: - It is defined as the conductivity of a solution per unit concentration

$$\text{i.e., } \Lambda_M = \frac{\kappa}{C}$$

It is clear from the above mathematical expression of the molar conductivity that, if we dilute the solution or decrease its concentration then molar conductivity increases. This is because, on dilution of a solution, a decrease in κ is more than compensated by the increase in its volume.

Question 3.8 The conductivity of 0.20 M solution of KCl at 298 K is 0.0248 S cm⁻¹. Calculate its molar conductivity.

Answer :

We know that the molar conductivity of a solution is defined as:-

$$\Lambda_M = \frac{\kappa}{C}$$

Putting the value of conductivity and concentration in the above equation:-

$$\Lambda_M = \frac{0.0248 \times 1000}{0.20} = 124 \text{ Scm}^2 \text{ mol}^{-1}$$

Question 3.9 The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500 ohm . What is the cell constant if conductivity of 0.001M KCl solution at 298 K is $0.146 \times 10^{-3} \text{ Scm}^{-1}$.

Answer:

We are given with conductivity of cell $\kappa = 0.146 \times 10^{-3} \text{ Scm}^{-1}$ and resistance R = 1500 Ω .

Also, **Cell constant** = $\kappa \times R$

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

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

Question 3.11 Conductivity of 0.00241 M acetic acid is $7.896 \times 10^{-5} \text{ Scm}^{-1}$.

Calculate its molar conductivity. If Δ_m^0 for acetic acid is $390.5 \text{ Scm}^2 \text{mol}^{-1}$, what is its dissociation constant?

Answer :

Molar conductivity of a solution is given by :-

$$\Lambda_M = \frac{\kappa}{C}$$

$$\text{So,} = \frac{7.896 \times 10^{-5}}{0.00241} \times 1000$$

$$\text{or} = 32.76 \text{ Scm}^2 \text{mol}^{-1}$$

Also, it is given that $\Lambda_m^0 = 390.5 \text{ Scm}^2 \text{mol}^{-1}$.

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

$$\text{or } \alpha = \frac{32.76}{390.5}$$

$$\alpha = 0.084$$

For dissociation constant we have,

$$K_d = \frac{c\alpha^2}{(1-\alpha)}$$

$$\text{so, } = \frac{0.00241 \times 0.084^2}{(1-0.084)}$$

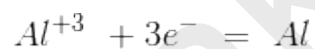
$$\text{or } = 1.86 \times 10^{-5} \text{ mol } L^{-1}$$

Question 3.12(i) How much charge is required for the following reductions:

(i) 1 mol of Al^{3+} to Al ?

Answer :

The equation becomes :-



So required charge is 3F.

$$Q = n \times 96500$$

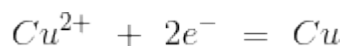
$$Q = 3 \times 96500 = 289500 \text{ C}$$

Question 3.12(ii) How much charge is required for the following reductions:

(ii) 1 mol of Cu^{2+} to Cu ?

Answer :

The equation can be written as:-



Thus charge required is = $2F$

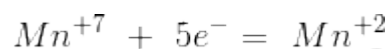
$$= 2(96500) = 193000 \text{ C}$$

Question 3.12(iii) How much charge is required for the following reductions:

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

Answer :

The given reaction can be written as:-



Thus charge required in above equation = $5F$

$$= 5(96500)$$

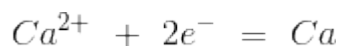
$$= 482500 \text{ C}$$

Question 3.13(i) How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten $CaCl_2$?

Answer :

The equation for the question is given by :-



In this equation, for 1 mol of Ca, 2F charge is required or we can say that for 40 g of Ca charge required is 2F.

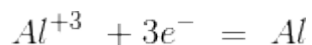
So, for 20 g of Ca charge required will be = F = 96500 C.

Question 3.13(ii) How much electricity in terms of Faraday is required to produce

(ii) 40.0 g of Al from molten Al_2O_3 ?

Answer :

The equation for the given question is :-



Thus for 1 mol of Al, charge required is 3F.

So the required amount of electricity in terms of charge will be :-

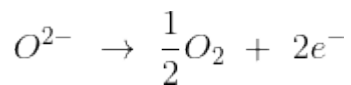
$$= \frac{3}{27} \times 40F = 4.44F$$

Question 3.14(i) How much electricity is required in coulomb for the oxidation of

(i) 1 mol of H_2O to O_2 ?

Answer :

According to question the equation of oxidation will be :-



Thus, for oxidation of O^{2-} , 2F charge is required.

$$= 2 \times 96500 \text{ C}$$

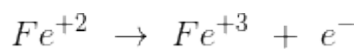
$$= 193000 \text{ C}$$

Question 3.14(ii) How much electricity is required in coulomb for the oxidation of

(ii) 1 mol of FeO to Fe_2O_3 ?

Answer :

The oxidation equation for the given reaction will be :-



So for oxidation of 1 mol Fe^{+2} charge required = 1F

$$= 96500 \text{ C}$$

Question 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 :

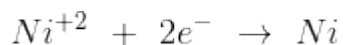
We are given:

$$I = 5A$$

$$\text{and } t = 20(60) = 1200 \text{ sec.}$$

So total charge = $5(1200) = 6000 \text{ C}$.

The equation for nickel deposition will be:-



Thus, from 2F charge 58.7 g of nickel deposition takes place.

$$\text{i.e., } 2(96487) \text{ C} \rightarrow 58.7 \text{ g Ni}$$

So for 6000 C charge total nickel deposition will be:-

$$= \frac{58.7}{2 \times 96487} \times 6000$$

$$\text{or} = 1.825 \text{ g}$$

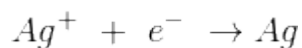
Hence 1.825 g Ni will be deposited in the given conditions.

Question 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 1.45g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

Answer :

Since the cells are connected in series so the current passing through each cell will be equal.(1.5 A)

Now we are given that 1.45 g of silver is deposited. So firstly we will consider the cell containing silver.



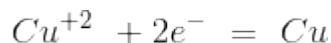
Since for deposition of 108 g silver 96487 C charge is required, thus for 1.45 g deposition of silver charge required will be:-

$$= \frac{96487}{108} \times 1.45 = 1295.43 \text{ C}$$

Now we can find the time taken by 1.5 A current to deposit 1.45 g silver.

$$\text{Time taken} = \frac{1295.43}{1.5} \approx 864 \text{ sec.}$$

For copper:-



Since 2F charge will deposit 63.5 g of Cu, then deposition by 1295.43 C will be:-

$$= \frac{63.5}{2 \times 96487} \times 1295.43 = 0.426 \text{ g}$$

Hence 0.426 g of copper will be deposited.

For zinc:-

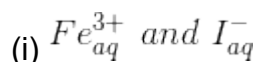


Since 2F charge will deposit 65.4 g of Zn, then deposition by 1295.43 C will be:-

$$= \frac{65.4}{2 \times 96487} \times 1295.43 = 0.439 \text{ g}$$

Hence 0.439 g of zinc will be deposited.

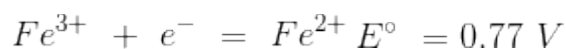
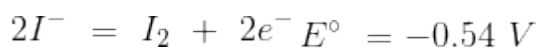
Question 3.17(i) Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:



Answer :

The concept used here will be that a reaction is feasible only if E_{cell}° is positive.

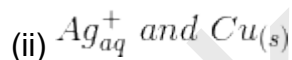
Anode and cathode reactions will be as follows:-



$$\text{So } E_{cell}^{\circ} = 0.77 - 0.54 = 0.23 \text{ V}$$

So this reaction is feasible.

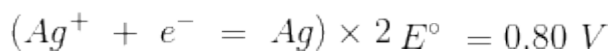
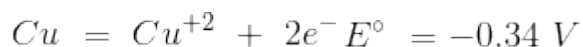
Question 3.17(ii) Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:



Answer :

A reaction is feasible only if E_{cell}° is positive.

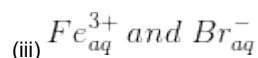
So, anode and cathode reactions will be as follows :-



and $E_{cell}^{\circ} = 0.80 - 0.34 = 0.46 \text{ V}$

So this reaction is feasible.

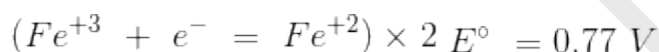
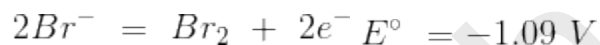
Question 3.17(iii) Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:



Answer :

A reaction is feasible only if E_{cell}° is positive.

So, anode and cathode reactions will be as follows :-



and $E_{cell}^{\circ} = 0.77 - 1.09 = -0.32 \text{ V}$

So this reaction is not feasible.

Question 3.17(iv) Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:



Answer :

A reaction is feasible only if E_{cell}° is positive.

So, anode and cathode reactions will be as follows:-

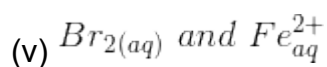
$$Ag = Ag^+ + e^- \quad E^\circ = -0.80 \text{ V}$$

$$Fe^{+3} + e^- = Fe^{+2} \quad E^\circ = 0.77 \text{ V}$$

$$\text{and } E_{cell}^\circ = -0.80 + 0.77 = -0.03 \text{ V}$$

So this reaction is not feasible.

Question 3.17(v) Using the standard electrode potentials given in Table 3.1, predict if the reaction between the following is feasible:



Answer :

A reaction is feasible only if E_{cell}° is positive.

So, anode and cathode reactions will be as follows :-

$$Br_2 + 2e^- = 2Br^- \quad E^\circ = 1.09 \text{ V}$$

$$Fe^{2+} = Fe^{3+} + e^- \quad E^\circ = -0.77 \text{ V}$$

$$\text{and } E_{cell}^\circ = 1.09 - 0.77 = 0.32 \text{ V}$$

So this reaction is feasible.

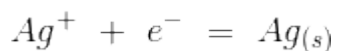
Question 3.18(i) Predict the products of electrolysis in each of the following:

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

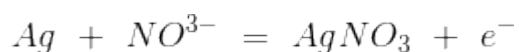
Answer :

For the given solution :

At cathode :- Reaction with greater E^0 will take place.



At anode :-



Hence, silver will get deposited at the cathode and it will be getting dissolved at anode.

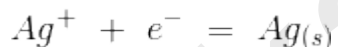
Question 3.18(ii) Predict the products of electrolysis in each of the following:

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

Answer :

For the given solution :

At cathode :- Reaction with greater E^0 will take place.



At anode :- Self ionisation will take place due to presence of water.



Hence, silver will get deposited at the cathode and O_2 will be produced from anode.

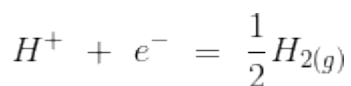
Question 3.18(iii) Predict the products of electrolysis in each of the following:

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

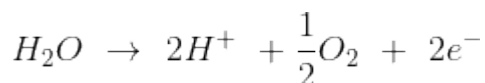
Answer :

For the given solution :

At cathode :- Reaction with greater E^0 will take place.



At anode :- Self ionisation of water will take place due to presence of platinum electrode.



Hence, H_2 gas will be generated at cathode and O_2 will be produced from anode.

Question 3.18(iv) Predict the products of electrolysis in each of the following:

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

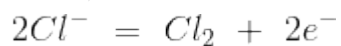
Answer :

For the given solution :

At cathode :- Reaction with greater E^0 will take place.



At anode :-



Hence, Cu will get deposited at cathode and Cl_2 will be produced from anode.