

EXERCISE

Q.1 Multiple choice questions. For each question there are four possible answer a, b, c and d. chose the one your consider correct.

- (i) The cathodic reaction in the electrolysis of dil. H_2SO_4 with Pt electrodes is:
- (a) Reduction (b) Oxidation
(c) Both oxidation and reduction (d) Neither oxidation or reduction
- (ii) Which of the following statements is not correct about galvanic cell?
- (a) Anode is negatively charged (b) Reduction occurs at anode
(c) Cathode is positively charge (d) Reduction occurs at cathode.
- (iii) Stronger the oxidizing agent, greater is the:
- (a) Oxidation potential (b) Reduction potential
(c) Redox potential (d) E.M.F of cell.
- (iv) If the salt bridge is not used between two half cells, then the voltage.
- (a) Decrease rapidly (b) Decrease slowly
(c) Does not change (d) Drops to Zero
- (v) If a strip of Cu metal is placed in a solution of FeSO_4 :
- (a) Cu will be precipitated out (b) Fe is precipitated out
(c) Cu and Fe both dissolve (d) No reaction take place

ANSWERS

Answers	Reasons
(i) (a)	When electrolysis of dil. H_2SO_4 is done with Pt-electrodes, H^+ ions are reduced to H_2 at cathode. $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$
(ii) (b)	Reduction always takes place at cathode and oxidation at anode. In galvanic cell, anode is negatively charged because loss of electrons takes place here whereas cathode is positively charged because electrons enter into the cell through it and reduction takes place here.
(iii) (b)	Reduction potential is the tendency of an element to gain electrons and to act as oxidizing agent. Greater the reduction potential. Stronger the oxidizing agent.
(iv) (d)	In the absence of salt bridge, there is net charge accumulation in the two half cells. As a result of that, the two half cell reactions stops and cell voltage drops to zero.
(v) (d)	Cu being below Fe in electrochemical series, cannot displace it from FeSO_4 . And hence, no. reaction occurs.



Q.2 Fill in the blanks:

- (i) The oxidation number of O-atom is _____ in OF_2 and is _____ in H_2O_2
- (ii) Conductivity of metallic conductors is due to the flow of _____ while that of electrolytes is due to flow of _____.
- (iii) Reaction taking place at the _____ is termed as oxidation and at the _____ is called as reduction.
- (iv) _____ is set up when a metal is dipped in its own ions.
- (v) Cu metal _____ the Cu-cathode when electrolysis is performed for CuSO_4 solution with Cu-cathodes.
- (vi) The reduction potential of Zn is _____ volts and its oxidation potential is _____ volts.
- (vii) In a fuel cell, _____ react together in the presence of _____.

ANSWERS

(i)	+2, -1	(ii)	Electrons, ions	(iii)	Anode, cathode
(iv)	Electrode potential	(v)	Deposited at	(vi)	-0.76, +0.76
(vii)	Hydrogen and oxygen, KOH				

Q.3 Mark the following statement true or false:

- (i) In electrolytic conduction, electrons flow through the electrolyte.
- (ii) In the process of electrolysis, the electrons in the external circuit flow from cathode to anode.
- (iii) Sugar is a non-electrolyte in solid form and when dissolved in water will allow the passage of an electric current.
- (iv) A metal will only allow the passage of an electric current when it is in cold state.
- (v) The electrolytic products of aqueous copper (II) chloride solution are copper and chlorine.
- (vi) Zinc can displace iron from its solution.
- (vii) S.H.E acts as cathode when connected with Cu-electrode.
- (viii) A voltaic cell produces electrical energy at the expense of chemical energy.
- (ix) Lead storage battery is not a reversible cell.
- (x) Cr changes its oxidation number when $\text{K}_2\text{Cr}_2\text{O}_7$ is reacted with HCl.



ANSWERS

Answers	Reasons
(i) False	In electrolytic conduction, ions flow through the electrolyte.
(ii) False	In the process of electrolysis, the electrons in the external circuit flow from anode to cathode.
(iii) False	Sugar is non-electrolyte in solid form and when dissolved in water will not allow the current to pass through it.
(iv) False	Metal will allow the passage of electric current when it is in cold or hot state.
(v) True	In case of aqueous solution of CuCl_2 , Cu^{2+} having greater reduction potential, in comparison to H^+ is discharge at cathode to form copper. Whereas at anode, Cl^- ions are discharged in preference with OH^- and Cl_2 gas is formed.
(vi) True	Since zinc lies above iron in the electrochemical series. Therefore, it can displace iron from its solution.
(vii) False	S.H.E acts as anode when connected with Cu-electrode. (Because copper has 0.34 volt more tendency to accept electrons than hydrogen and hence copper act as cathode).
(viii) True	In voltaic cell, spontaneous redox reaction proceeds to produce electrical energy.
(ix) False	Lead storage battery is a reversible cell.
(x) True	When $\text{K}_2\text{Cr}_2\text{O}_7$ is reacted with HCl , Cr changes its oxidation state from +6 to +3.

- Q.4**
- (a) Explain the term oxidation number with examples.
- (b) Describe the rules used for the calculation of oxidation number of an element in molecules and ions giving examples.
- (c) Calculate the oxidation number of chromium in the following compounds:
- (i) CrCl_3 (ii) $\text{Cr}_2(\text{SO}_4)_3$ (iii) K_2CrO_4
 (iv) $\text{K}_2\text{Cr}_2\text{O}_7$ (v) CrO_3 (vi) Cr_2O_3
 (vii) $\text{Cr}_2\text{O}_7^{-2}$
- (d) Calculate the oxidation number of elements underlined in the following compounds.
- (i) $\text{Ca}(\underline{\text{ClO}}_3)_2$ (ii) $\text{Na}_3\underline{\text{P}}\text{O}_4$ (iii) $\text{Cr}_2(\underline{\text{S}}\text{O}_4)_3$
 (iv) $\text{K}_2\underline{\text{Mn}}\text{O}_4$ (v) $\text{Na}_2\underline{\text{C}}\text{O}_3$ (vi) $\text{H}\underline{\text{N}}\text{O}_3$
 (vii) $\text{H}\underline{\text{P}}\text{O}_3$

Ans. For answer of part “a” and “b”, see text book.

(c) (i) CrCl_3 :

$$\text{Oxidation no. of Cr} + (\text{Oxidation no. of Cl})_3 = 0$$

$$\text{Cr} + 3(-1) = 0$$

$$\text{Cr} = +3$$

(ii) $\text{Cr}_2(\text{SO}_4)_3$:

$$2(\text{O.N of Cr}) + 3(\text{O.N of S}) + 12(\text{O.N of O}) = 0$$

$$2\text{Cr} + 3(+6) + 12(-2) = 0$$

$$2\text{Cr} + 18 + (-24) = 0$$

$$2\text{Cr} + 18 - 24 = 0$$

$$2\text{Cr} = 6$$

$$\text{Cr} = \frac{6}{2} = +3$$

(iii) K_2CrO_4 :

$$2(\text{O.N of K}) + \text{O.N of Cr} + 4(\text{O.N of O}) = 0$$

$$2(+1) + \text{Cr} + 4(-2) = 0$$

$$2 + \text{Cr} - 8 = 0$$

$$\text{Cr} = +6$$

(iv) $\text{K}_2\text{Cr}_2\text{O}_7$:

$$2(\text{O.N of K}) + 2(\text{O.N of Cr}) + 7(\text{O.N of O}) = 0$$

$$2(+1) + 2\text{Cr} + 7(-2) = 0$$

$$+2 + 2\text{Cr} - 14 = 0$$

$$2\text{Cr} = 12$$

$$\text{Cr} = \frac{12}{2} = +6$$

(v) CrO_3 :

$$\text{O.N of Cr} + 3(\text{O.N of O}) = 0$$

$$\text{Cr} + 3(-2) = 0$$

$$\text{Cr} - 6 = 0$$

$$\text{Cr} = +6$$

(vi) Cr_2O_3 :

$$2(\text{O.N of Cr}) + 3(\text{O.N of O}) = 0$$

$$2\text{Cr} + 3(-2) = 0$$

$$2\text{Cr} - 6 = 0$$

$$2\text{Cr} = +6$$

$$\text{Cr} = \frac{6}{2} = +3$$

(vii) $\text{Cr}_2\text{O}_7^{2-}$:

$$2(\text{O.N of Cr}) + 7(-2) = -2$$

$$2\text{Cr} - 14 = -2$$

$$2\text{Cr} = 14 - 2$$

$$\text{Cr} = \frac{12}{2} = +6$$



(d) (i) $\text{Ca}(\text{ClO}_3)_2$:

$$\begin{aligned}
 \text{O.N of Ca} + (\text{O.N of Cl})_2 + 6(\text{O.N of O}) &= 0 \\
 +2 + 2\text{Cl} + 6(-2) &= 0 \\
 +2 + 2\text{Cl} - 12 &= 0 \\
 2\text{Cl} &= 10 \\
 \text{Cl} &= \frac{10}{2} = +5
 \end{aligned}$$

(ii) Na_3PO_4 :

$$\begin{aligned}
 3(\text{O.N of Na}) + \text{O.N of P} + 4(\text{O.N of O}) &= 0 \\
 3(+1) + \text{P} + 4(-2) &= 0 \\
 3 + \text{P} - 8 &= 0 \\
 \text{P} &= +5
 \end{aligned}$$

(iii) $\text{Cr}_2(\text{SO}_4)_3$:

$$\begin{aligned}
 2(\text{O.N of Cr}) + 3(\text{O.N of S}) + 12(\text{O.N of O}) &= 0 \\
 2(+3) + 3\text{S} + 12(-2) &= 0 \\
 6 + 3\text{S} - 24 &= 0 \\
 3\text{S} &= 18 \\
 \text{S} &= \frac{18}{3} = +6
 \end{aligned}$$

(iv) K_2MnO_4 :

$$\begin{aligned}
 2(\text{O.N of K}) + \text{O.N of Mn} + 4(\text{O.N of O}) &= 0 \\
 2(+1) + \text{Mn} + 4(-2) &= 0 \\
 2 + \text{Mn} - 8 &= 0 \\
 \text{Mn} &= +6
 \end{aligned}$$

(v) Na_2CO_3 :

$$\begin{aligned}
 2(\text{O.N of Na}) + \text{O.N of C} + 3(\text{O.N of O}) &= 0 \\
 2(+1) + \text{C} + 3(-2) &= 0 \\
 \text{C} &= +4
 \end{aligned}$$

(vi) HNO_3 :

$$\begin{aligned}
 (\text{O.N of H}) + \text{O.N of N} + 3(\text{O.N of O}) &= 0 \\
 +1 + \text{N} + 3(-2) &= 0 \\
 1 + \text{N} - 6 &= 0 \\
 \text{N} &= +5
 \end{aligned}$$



- Q.5** (a) Describe the general rules for balancing a redox equation by oxidation number, method.
- (b) Balance the following equations by oxidation number method.
- $\text{Cu} + \text{HNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O}$
 - $\text{Zn} + \text{HNO}_3 \longrightarrow \text{Zn}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$
 - $\text{Br}_2 + \text{NaOH} \longrightarrow \text{NaBr} + \text{NaBrO}_3 + \text{H}_2\text{O}$
 - $\text{MnO}_2 + \text{HCl} \longrightarrow \text{MnCl}_2 + \text{H}_2\text{O} + \text{Cl}_2$
 - $\text{FeSO}_4 + \text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{SO}_4 \longrightarrow \text{Fe}_2(\text{SO}_4)_3 + \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$
 - $\text{HNO}_2 + \text{HI} \longrightarrow \text{NO} + \text{H}_2\text{O} + \text{I}_2$
 - $\text{Cu} + \text{H}_2\text{SO}_4 \longrightarrow \text{CuSO}_4 + \text{SO}_2 + \text{H}_2\text{O}$
 - $\text{HI} + \text{H}_2\text{SO}_4 \longrightarrow \text{I}_2 + \text{SO}_2 + \text{H}_2\text{O} + \text{Cl}$
 - $\text{NaCl} + \text{H}_2\text{SO}_4 + \text{MnO}_2 \longrightarrow \text{Na}_2\text{SO}_4 + \text{MnSO}_4 + \text{H}_2\text{O} + \text{Cl}_2$

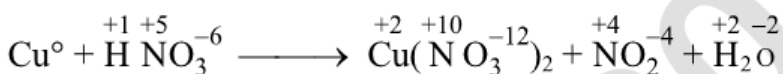
Ans.

(a) Descriptive question, consult text book.

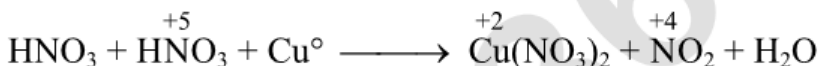
(b) (i) $\text{Cu} + \text{HNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O}$

Steps Involved:

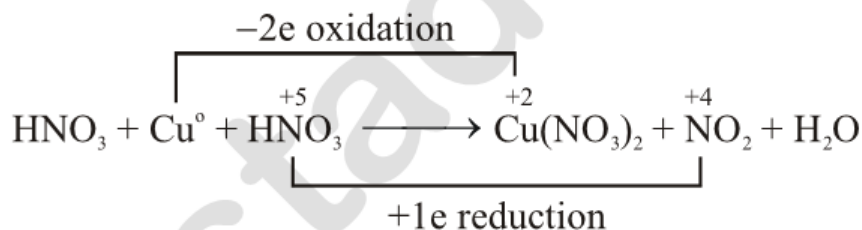
1. Write the oxidation number of each element.



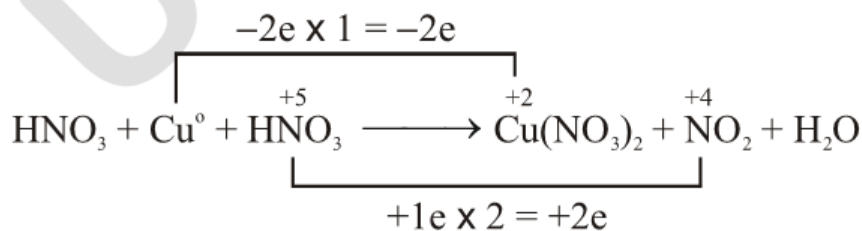
2. Identify the elements whose oxidation number have changed.



3. Mention the changed by arrow, also balance same atom on both sides, if they are not equal. Determine change in oxidation and reduction process by loss and gain of electron.



4. Equate the number of electrons by multiply -2e with 1 and +1e with 2. So that gain and loss of electrons become equal.



$$\text{HNO}_3 + \text{Cu} + 2\text{HNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + 2\text{NO}_2 + \text{H}_2\text{O}$$
$$2\text{HNO}_3 + \text{Cu} + 2\text{HNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + 2\text{NO}_2 + \text{H}_2\text{O}$$
$$\text{Cu} + 4\text{HNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + 2\text{NO}_2 + 2\text{H}_2\text{O}$$

* Steps Involved:

$$2. \quad \text{HNO}_3 + \text{Zn}^0 + \overset{+5}{\text{HNO}_3} \longrightarrow \overset{+2}{\text{Zn}(\text{NO}_3)_2} + \overset{+2}{\text{NO}} + \text{H}_2\text{O}$$

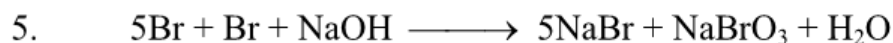
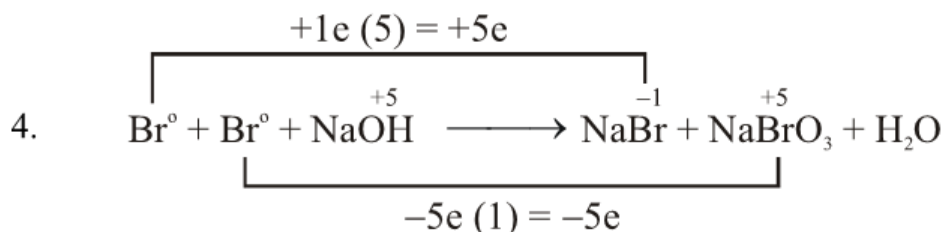
4.
$$\text{HNO}_3 + \text{Zn}^0 + \text{HNO}_3 \longrightarrow \text{Zn}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$$

6. $6\text{HNO}_3 + 3\text{Zn} + 2\text{HNO}_3 \longrightarrow 3\text{Zn}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$
$$3\text{Zn} + 8\text{HNO}_3 \longrightarrow 3\text{Zn}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$$

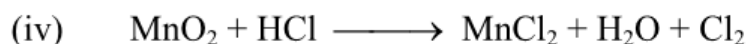
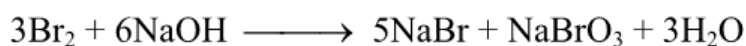
* Steps Involved:

$$2. \quad \text{Br}^0 + \text{Br}^0 + \text{NaOH} \longrightarrow \overset{-1}{\text{NaBr}} + \overset{+5}{\text{NaBrO}_3} + \text{H}_2\text{O}$$

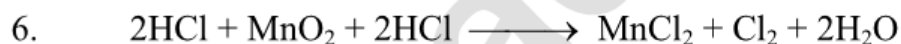
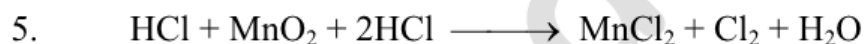
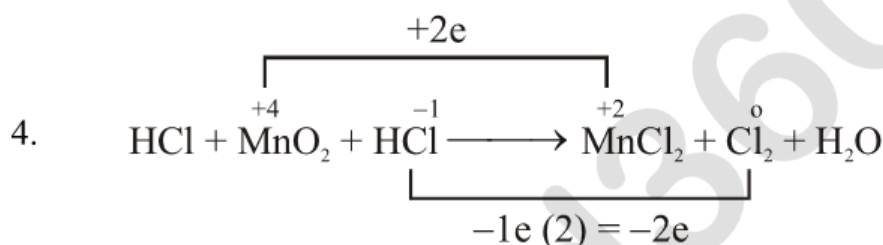
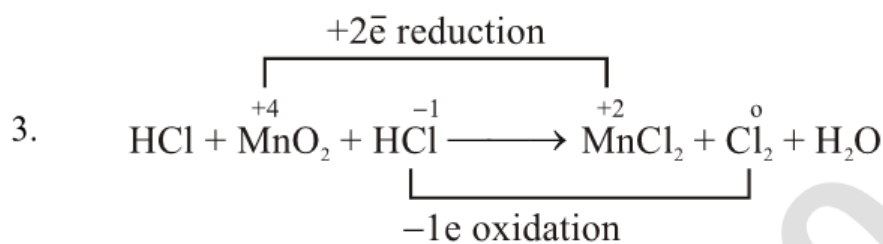
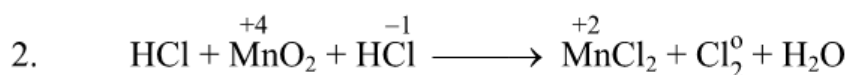
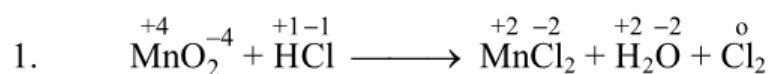
3.
$$\text{Br}^0 + \text{Br}^0 + \overset{+5}{\text{NaOH}} \longrightarrow \overset{-1}{\text{NaBr}} + \overset{+5}{\text{NaBrO}_3} + \text{H}_2\text{O}$$



* Balanced Equation:



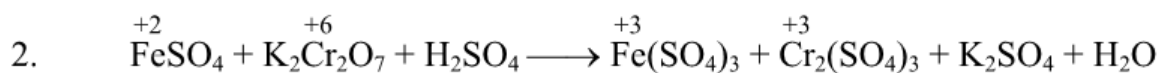
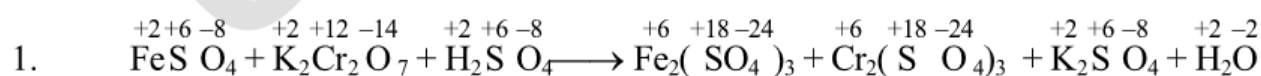
* Steps Involved:

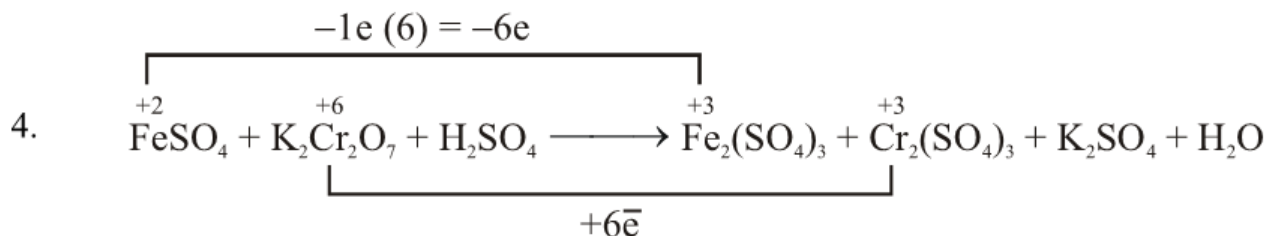
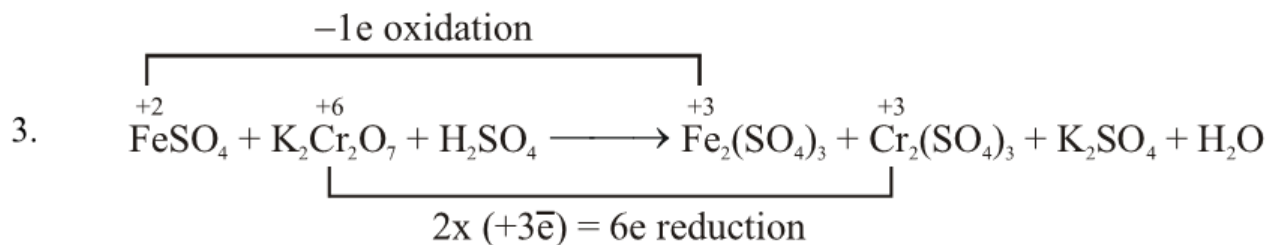


* Balanced Equation:

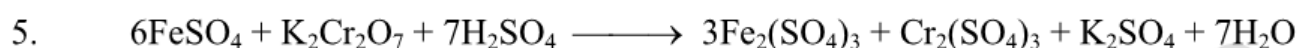


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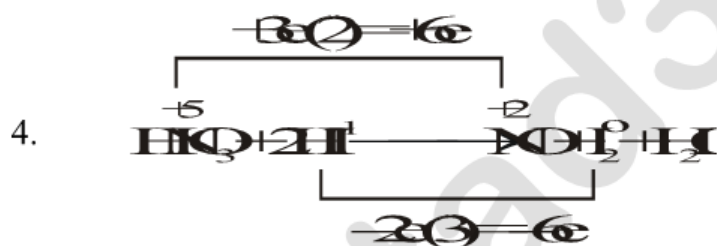
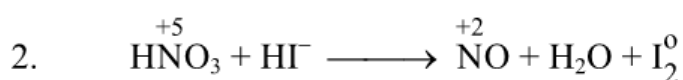
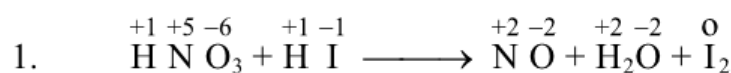




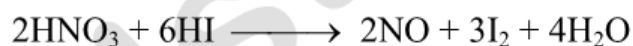
* Balanced Equation:



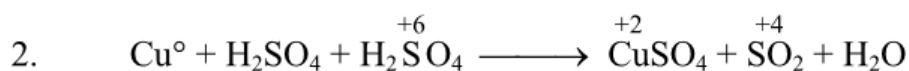
* Steps Involved:

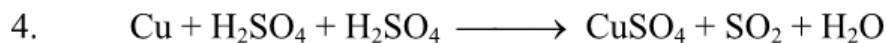
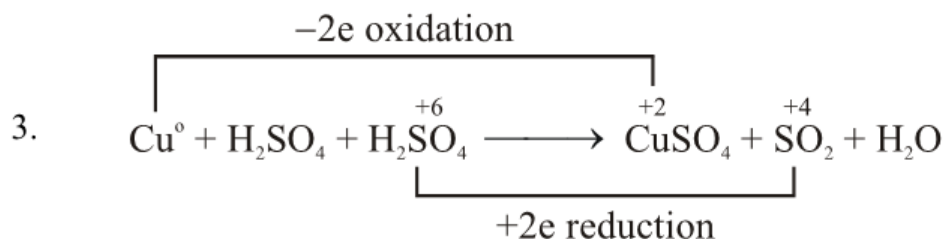


* Balanced Equation:

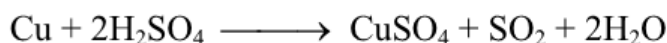


* Steps Involved:

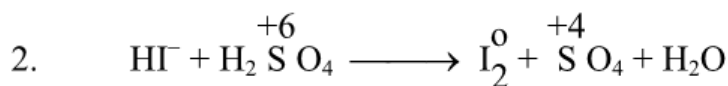
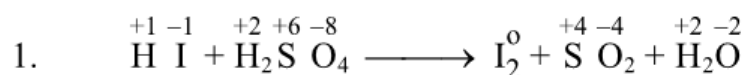




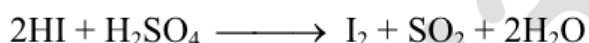
* Balanced Equation:



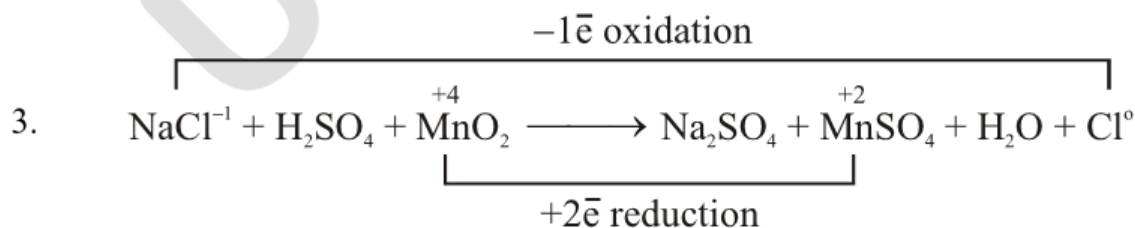
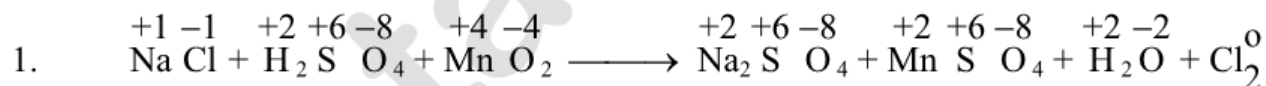
* Steps Involved:



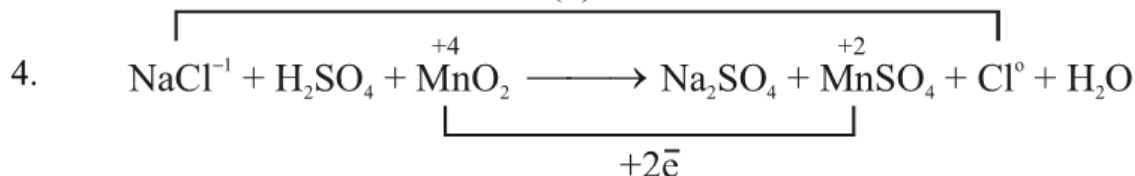
* Balance Equation:



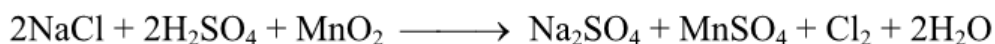
* Steps Involved:



$$-1\bar{e} (2) = -2e$$



* Balanced Equation:



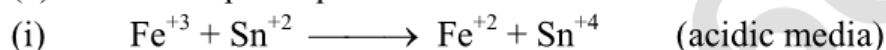
Q.6 (a) Describe the general rules of balancing a redox equation by ion-electron method.

(b) Balance the following ionic equations by ion-electron method.

- (i) $\text{Fe}^{+3} + \text{Sn}^{+2} \longrightarrow \text{Fe}^{+2} + \text{Sn}^{+4}$
- (ii) $\text{MnO}_4^{-1}(\text{aq}) + \text{C}_2\text{O}_4^{-2} \longrightarrow \text{Mn}^{+2}(\text{aq}) + \text{CO}_2(\text{g})$
- (iii) $\text{Cr}_2\text{O}_7^{-2} + \text{Cl}^{-} \longrightarrow \text{Cr}^{+3} + \text{Cl}_2$
- (iv) $\text{Cu} + \text{NO}_3^{-1} \longrightarrow \text{Cu}^{+2} + \text{NO}_2$
- (v) $\text{Cr}_2\text{O}_7^{-2} + \text{Fe}^{+2} \longrightarrow \text{Cr}^{+3} + \text{Fe}^{+3}$ (acidic media)
- (vi) $\text{S}_2\text{O}_3^{-2} + \text{OCl}^{-1} \longrightarrow \text{Cl}^{-} + \text{S}_4\text{O}_6^{-2}$ (acidic media)
- (vii) $\text{IO}_3^{-1} + \text{AsO}_3^{-3} \longrightarrow \text{I} + \text{AsO}_4^{-3}$ (acidic media)
- (viii) $\text{Cr}^{+3} + \text{BiO}_3^{-1} \longrightarrow \text{Cr}_2\text{O}_7^{-2} + \text{Bi}^{+3}$ (acidic media)
- (ix) $\text{H}_2\text{AsO}_3 + \text{Cr}_2\text{O}_7^{-2} \longrightarrow \text{H}_3\text{AsO}_4 + \text{Cr}^{+3}$ (acidic media)
- (x) $\text{CN}^{-} + \text{MnO}_4^{-1} \longrightarrow \text{C}\bar{\text{N}}\text{O} + \text{MnO}_2(\text{s})$ (basic media)

Ans.

(a) Descriptive question consult text book for details.



* Steps Involved:

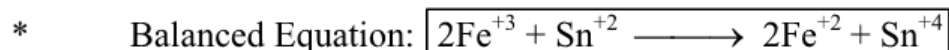
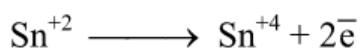
1. Split the equation into two halves.

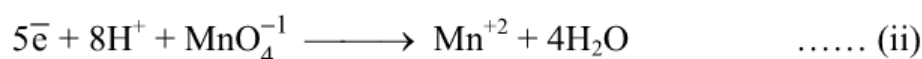
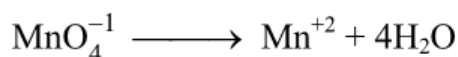
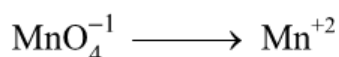
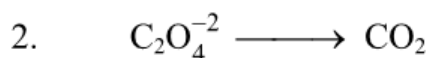
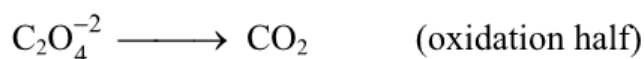
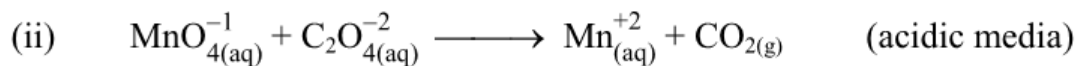


2. Balance the atoms and changes on both the sides of half reactions separately.

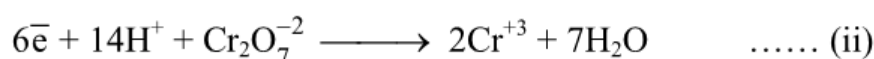
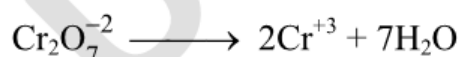
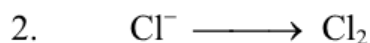
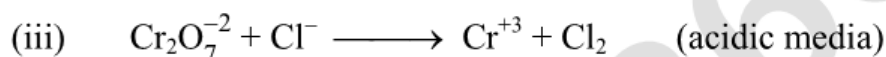
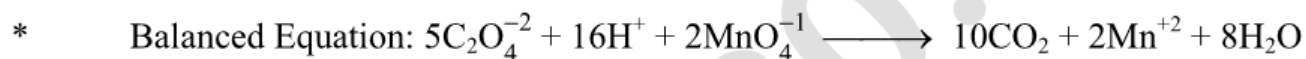
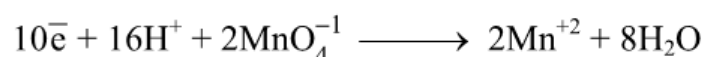
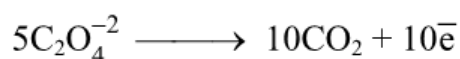


3. Multiply eq. (ii) by 2 and add it into eq. (i):

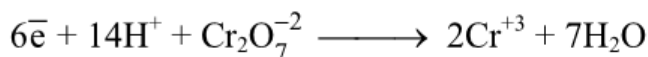
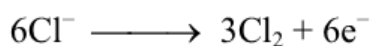




3. Multiply eq. (i) by 5 and eq. (ii) by 2 and then add:



3. Multiply eq. (i) by 3 and add it into eq. (ii):



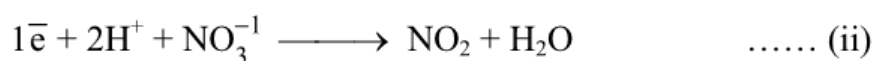
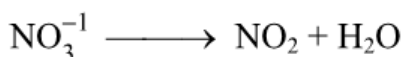
* Balance Equation: $6\text{Cl}^- + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} \longrightarrow 3\text{Cl}_2 + 2\text{Cr}^{+3} + 7\text{H}_2\text{O}$

(iv) $\text{Cu} + \text{NO}_3^{-1} \longrightarrow \text{Cu}^{+2} + 2\text{NO}_2$ (acidic media)

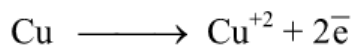
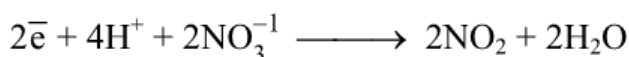
1. $\text{Cu} \longrightarrow \text{Cu}^{+2}$ (oxidation half)

$\text{NO}_3^{-1} \longrightarrow \text{NO}_2$ (reduction half)

2. $\text{Cu} \longrightarrow \text{Cu}^{+2} + 2\text{e}^-$ (i)



3. Multiply eq. (ii) by 2 and add it into (i):



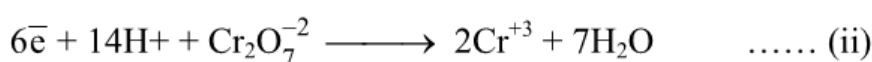
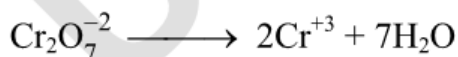
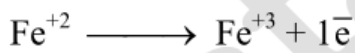
* Balance Equation: $4\text{H}^+ + 2\text{NO}_3^{-1} + \text{Cu} \longrightarrow 2\text{NO}_2 + 2\text{H}_2\text{O} + \text{Cu}^{+2}$

(v) $\text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{+2} \longrightarrow \text{Cr}^{+3} + \text{Fe}^{+3}$ (acidic medium)

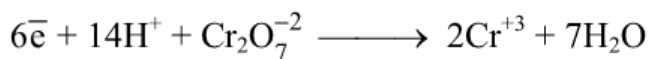
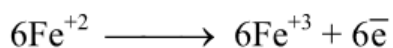
1. $\text{Cr}_2\text{O}_7^{2-} \longrightarrow \text{Cr}^{+3}$ (reduction half)

$\text{Fe}^{+2} \longrightarrow \text{Fe}^{+3}$ (oxidation half)

2. $\text{Fe}^{+2} \longrightarrow \text{Fe}^{+3}$ (i)



3. Multiply eq. (i) by 6 and add it into eq. (ii):



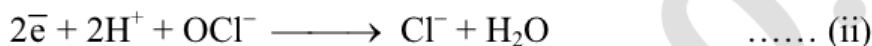
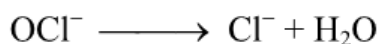
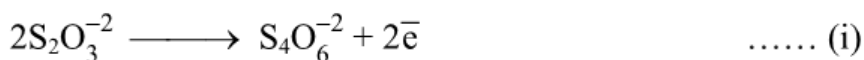
* Balance Equation: $6\text{Fe}^{+2} + 14\text{H}^+ + \text{Cr}_2\text{O}_7^{-2} \longrightarrow 6\text{Fe}^{+3} + 2\text{Cr}^{+3} + 7\text{H}_2\text{O}$

(vi) $\text{S}_2\text{O}_3^{-2} + \text{OCl}^- \longrightarrow \text{S}_4\text{O}_6^{-2} + \text{Cl}^-$ (acidic medium)

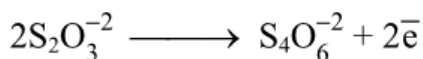
1. $\text{S}_2\text{O}_3^{-2} \longrightarrow \text{S}_4\text{O}_6^{-2}$ (oxidation half)

$\text{OCl}^- \longrightarrow \text{Cl}^-$ (reduction half)

2. $\text{S}_2\text{O}_3^{-2} \longrightarrow \text{S}_4\text{O}_6^{-2}$



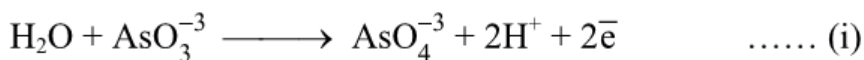
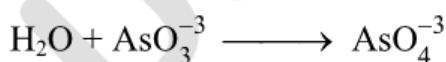
3. Add eq. (i) into (ii):

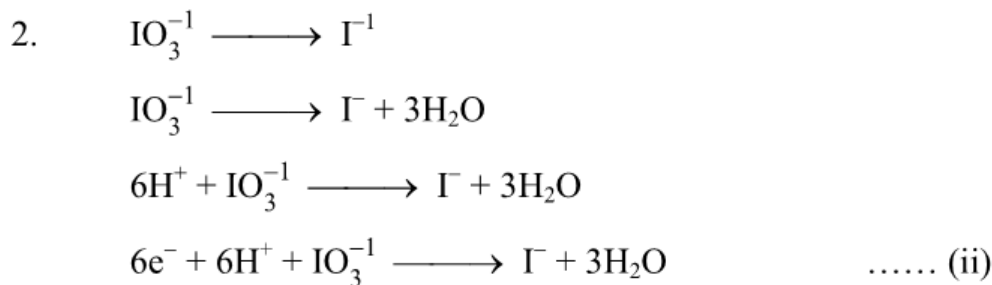


* Balance Equation: $2\text{S}_2\text{O}_3^{-2} + 2\text{H}^+ + \text{OCl}^- \longrightarrow \text{S}_4\text{O}_6^{-2} + \text{Cl}^- + \text{H}_2\text{O}$

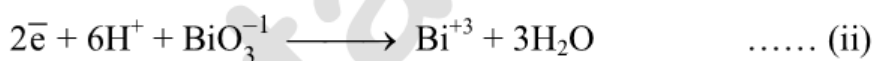
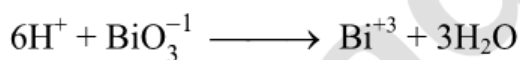
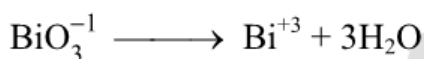
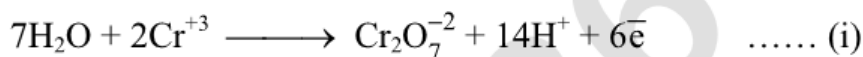
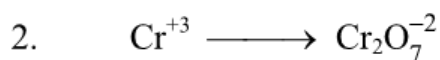
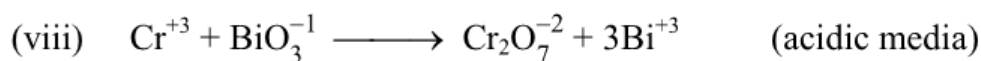
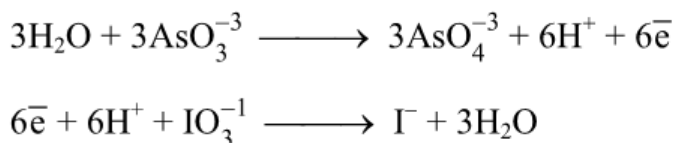
(vii) $\text{IO}_3^{-1} + \text{AsO}_3^{-3} \longrightarrow \text{I}^- + \text{AsO}_4^{-3}$ (acidic media)

1. $\text{AsO}_3^{-3} \longrightarrow \text{AsO}_4^{-3}$ (oxidation half)

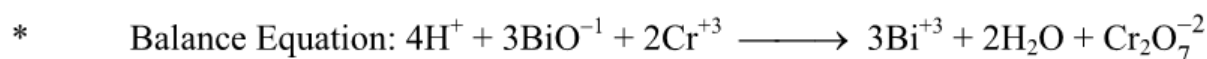
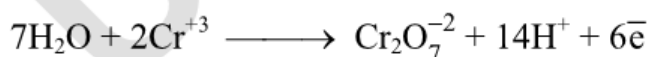
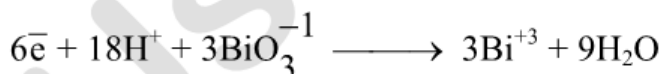


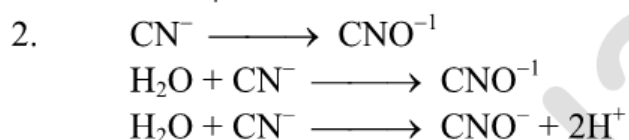
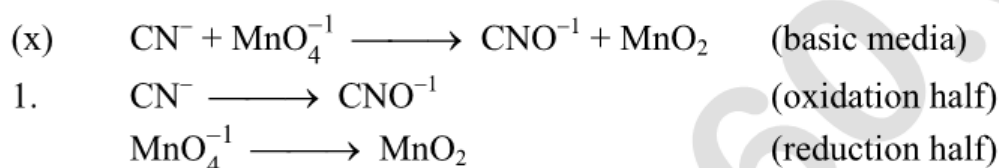
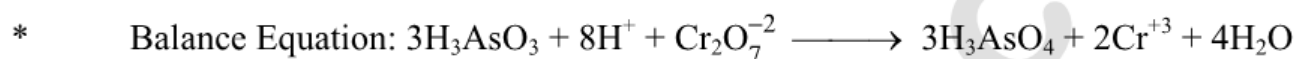
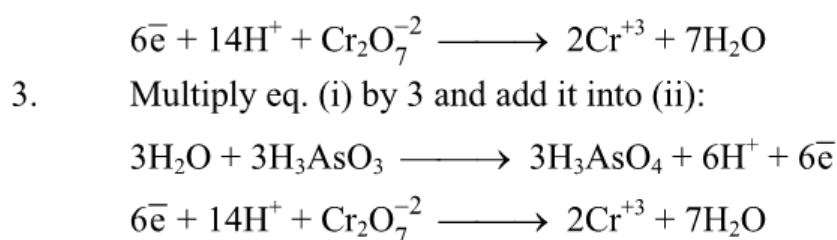
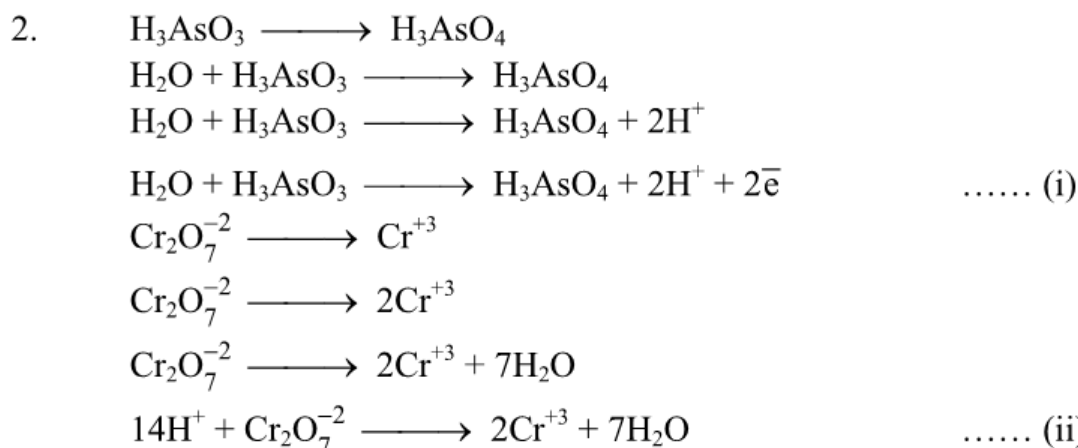
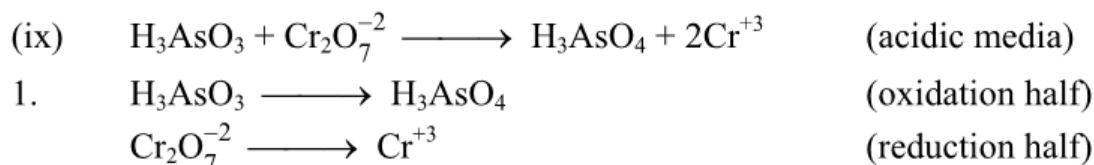


3. Multiply eq. (i) by 3 and add it into eq. (ii):

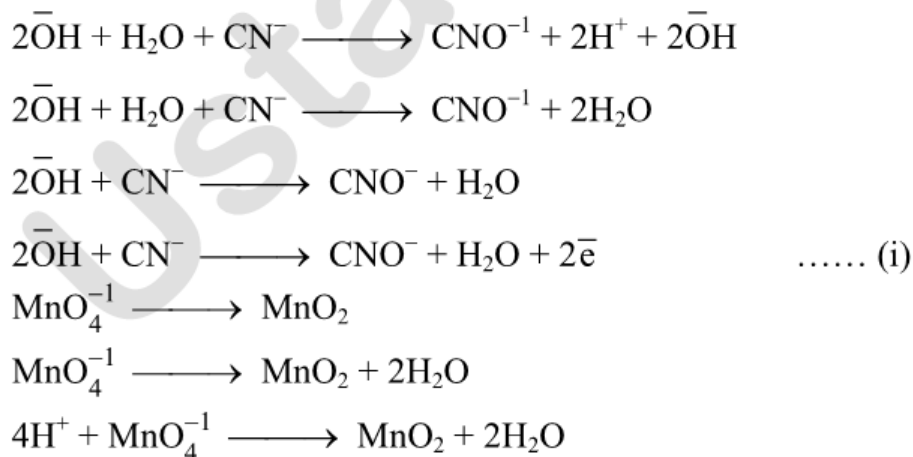


3. Multiply eq. (ii) by 3 and add it into eq. (i):

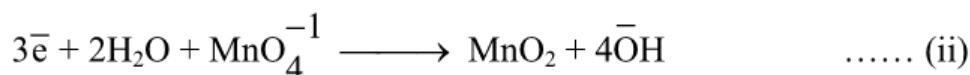
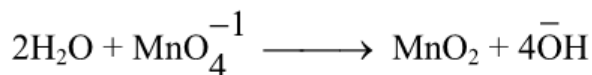
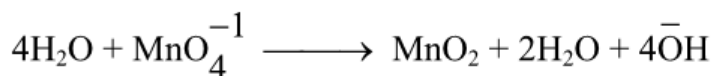




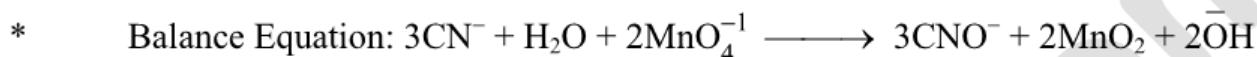
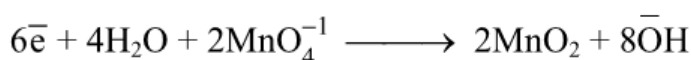
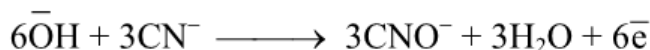
Since this is an equation of basic medium, add OH^- ions on both sides of equation equal to the no. of H^+ ions.



Since this is an equation of basic medium. Hence, we add OH^- ions equal to the no. of H^+ ions on both sides of equation.

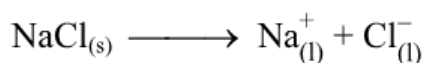


3. Multiply eq. (i) by 3 and eq. (ii) by 2 and add them.

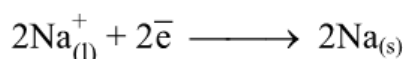


Q.7 Describe the electrolysis of molten sodium chloride, and a concentrated solution of sodium chloride.

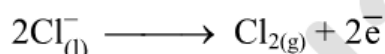
Ans. Electrolysis of Molten NaCl: Solid NaCl when heated, it melts and ionizes as following:



At Cathode:



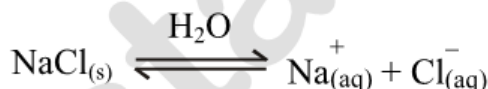
At Anode:



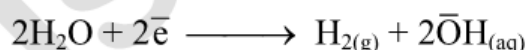
Overall Reaction:



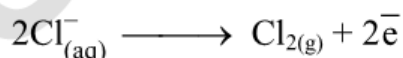
“Electrolysis of Aqueous” Solution of NaCl:



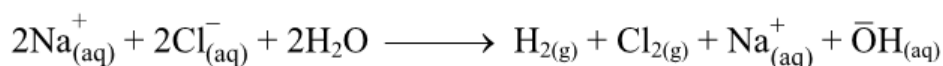
At Cathode:



At Anode:



Over Reaction:



Q.8 What is the difference between single electrode potential and standard electrode potential? How can it be measured? Give its importance.

Ans.

Single Electrode Potential	Standard Electrode Potential
The potential set up, when an electrode is in contact with one molar solution of its ions is called as single electrode potential.	The potential set up, when electrode is in contact with one molar solution of its own ions at 298K is called as standard electrode potential.

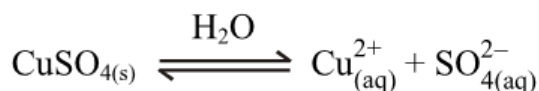
For measurement and importance consult text book.

Q.9 Outline the important applications of electrolysis. Write the electrochemical reactions involved there in. Discuss the electrolysis of CuSO_4 using Cu-electrodes and AgNO_3 solution using Ag electrode.

Ans.

For applications of electrolysis consult text book.

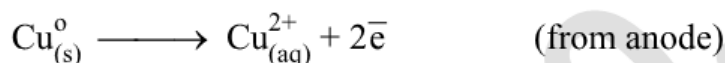
Electrolysis of CuSO_4 using Cu-electrode:



At Cathode:

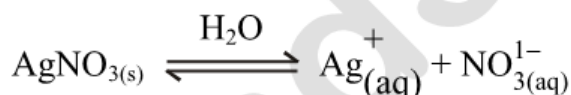


At Anode:

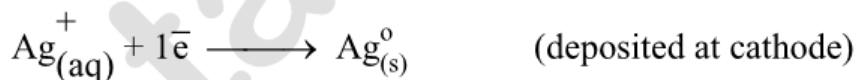


In this electrolysis CuSO_4 ionizes to produce Cu^{2+} ions and SO_4^{2-} . Cu^{2+} ion pick up electrons and are deposited at cathode. From anode, copper atoms loose electrons and form Cu^{2+} ions in the solution. In other words Cu from Cu-anode dissolves into the solution and is deposited at Cu-cathode.

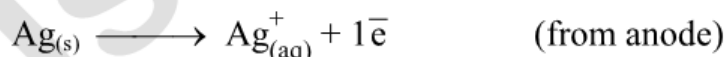
Electrolysis of AgNO_3 using Ag-electrode:



At Cathode:



At Anode:



In this type of electrolysis, AgNO_3 ionizes to give Ag^+ ions and NO_3^{1-} ions. Ag^+ ion pick up electrons and are deposited at cathode. From anode, $\text{Ag}_{(s)}$ atoms loose electrons to form Ag^+ ions in the solution.

In other words, Ag from Ag-anode dissolves into the solution and is deposited at Ag-cathode.



Q.10 Describe the construction and working of standard hydrogen electrode.

Ans. Descriptive question consult text book for details.

Q.11 Is the reaction $\text{Fe}^{3+} + \text{Ag} \rightarrow \text{Fe}^{2+} + \text{Ag}^+$ spontaneous? If not, write spontaneous reaction involving these species.

Ans. $\text{Fe}^{3+} + \text{Ag}^{\circ} \longrightarrow \text{Fe}^{2+} + \text{Ag}^+$

Reduction potential of Fe = +0.771 volt

Reduction potential of Ag = +0.799 volt

Since, Ag is oxidized in this reaction, its oxidation potential will be -0.799 volt.

$$E_{\text{cell}}^{\circ} = E_{\text{oxi}}^{\circ} + E_{\text{red}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = -0.799 + 0.771$$

$$E_{\text{cell}}^{\circ} = -0.028 \text{ volt}$$

Thus, the reaction is not spontaneous because, E° for the cell is negative. The reverse of above reaction is spontaneous and feasible. i.e.,



$$E_{\text{cell}}^{\circ} = E_{\text{oxi}}^{\circ} + E_{\text{red}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = -0.771 + 0.799$$

$$E_{\text{cell}}^{\circ} = +0.028 \text{ volt}$$

Q.12 Explain the difference between:

- Ionization and electrolysis.**
- Electrolytic cell and voltaic cell.**
- Conduction through metals and molten electrolytes.**

Ans.

(a) Difference between ionization and electrolysis:

Ionization	Electrolysis
(i) Splitting up of ionic compound into ions in molten state or in aqueous form is called ionization.	(i) Decomposition of a substance in solution or molten state by passing electricity is called electrolysis.
(ii) In this case, electrodes, electricity and electrolytic cell is not required.	(ii) For electrolysis electrodes, electricity and electrolytic cell is required.
(iii) During ionization, ions are produced.	(iii) During electrolysis, ions are oxidized or reduced to neutral atoms or molecules.
(iv) Example: $\text{NaCl}_{(s)} \xrightarrow{\text{H}_2\text{O}} \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}$	(iv) Example: $2\text{NaCl}_{(l)} \xrightarrow[\text{H}_2\text{O}]{\text{Electrolysis}} 2\text{Na}_{(s)} + \text{Cl}_{2(g)}$



(b) Difference between electrolytic cell and voltaic cell:

Electrolytic Cell	Voltaic Cell
(i) A cell in which electrical energy is converted into chemical energy is called electrolytic cell.	(i) A cell in which chemical energy is converted into electrical energy is called voltaic cell.
(ii) In this cell current is used.	(ii) In this cell current is produced.
(iii) Non-spontaneous redox reaction take place here.	(iii) Spontaneous redox reaction take place in it.
(iv) Example: Down's cell, Nelson's cell.	(iv) Example: Daniel cell, Fuel cell.

(c) Difference between conduction through metals and molten electrolytes:

Conduction Through Metals	Conduction Through Molten Electrolytes
(i) In metals, conduction takes place due to free electrons, so it is called as electronic conduction.	(i) In molten electrolytes, conduction takes place due to free ions, so it is called as electrolytic conduction.
(ii) In this case, conduction takes place through external circuit.	(ii) In this case, conduction takes place through liquid medium.
(iii) There is no change of properties of conductors.	(iii) There is decomposition of electrolytes as a result of chemical reaction during conduction.
(iv) There is increase of resistance with the rise of temperature.	(iv) There is decrease of resistance with rise of temperature.
(v) Examples: All metals.	(v) Examples: Molten salts.

Q.13 Describe a galvanic cell explaining the function of electrodes and the salt bridge.

Ans. Descriptive question consult text book.

Q.14 Write comprehensive notes on:

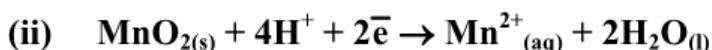
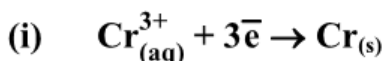
- Spontaneity of oxidation reduction reactions.
- Electrolytic conduction.
- Alkaline, silver oxide and nickel-cadmium batteries, fuel cell.
- Lead accumulator, its desirable and undesirable features.

Ans. Descriptive question consult text book.



Q.15 Will the reaction be spontaneous for the following set of half reactions.

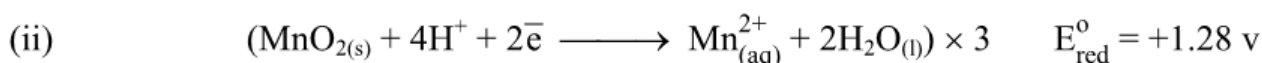
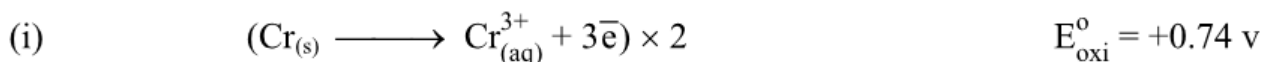
What will be the value of E° Cell?



Standard reduction potential for reaction.

(i) -0.74 v and for the reaction (ii) $= +1.28 \text{ V}$

Ans. Yes this reaction is spontaneous, when $\text{Cr}_{(\text{s})}$ with low reduction potential act as anode and MnO_2 with high reduction potential act as cathode:



$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{oxi}}^\circ + E_{\text{red}}^\circ \\ &= +0.74 + 1.28 \end{aligned}$$

$$E_{\text{cell}}^\circ = 2.02 \text{ volt}$$

Q.16 Explain the following with reason:

- A porous plate or a salt bridge is not required in lead storage cell.
- The standard oxidation potential of Zn is 0.76 V and its reduction potential is -0.76 V .
- Na and K can displace hydrogen from acids but Pt, Pd and Cu cannot.
- The equilibrium is set up between metal atoms of electrode and ions of metal in a cell.
- A salt bridge maintains the electrical neutrality in the cell.
- Lead accumulator is a chargeable battery.
- Impure Cu can be purified by electrolytic process.
- SHE acts as anode when connected with Cu electrode but as cathode with Zn electrode.

Ans.

- A porous plate or salt bridge is used in those cells where two different electrolytes are used and are required to be kept separate. In case of lead storage cell, only dil. H_2SO_4 is used as an electrolyte. Hence, no separation is required by porous plate or salt bridge.
- The tendency of an element measured when coupled with SHE is shown with the positive sign, and reverse tendency is shown with negative sign.
Zinc when coupled with SHE, has 0.76 v more tendency to show oxidation than hydrogen. Hence, its oxidation potential is 0.76 v and is shown with positive sign (+0.76 v). Therefore, its reduction potential will be shown with negative sign (-0.76 v) because it has 0.76 v less tendency to show reduction than hydrogen.



- (c) Na and K have very low value of reduction potential and lie very much above hydrogen in the electrochemical series. Hence, can displace hydrogen from acids.

Pt, Pd and Cu have very high value of reduction potentials and lie below hydrogen in electrochemical series. Therefore, these metals cannot displace hydrogen from acids.

- (d) When metal is dipped into the solution of its own ions, metal atoms lose electrons, form positive ions and enter into the solution. Already present ions in the solution, pick up electrons and are deposited on electrode.

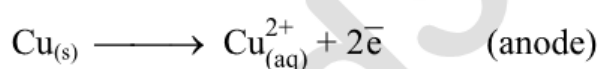
This process is continuous and a stage is reached when rate of both the processes becomes equal. This stage is called equilibrium stage.

- (e) A salt bridge is a U-tube containing saturated solutions of strong electrolytes. e.g., KCl, KNO_3 . It connects the two electrolytic solutions. When an electrochemical reaction takes place in the cell it allows the passage of negative ions from cathode compartment (where they are in excess) to anode compartment (where cations are in excess). As a result of that both compartments become neutral.

- (f) The lead accumulator is a secondary cell. It is charged by passing direct current through it. To recharge it, anode and cathode of external source are connected to its anode and cathode of lead accumulator respectively. The redox reactions at the respective electrodes are reversed during recharging again. The density of H_2SO_4 is also regained during this process.

- (g) During purification of Cu, impure Cu is made anode and thin sheet of pure Cu is made cathode and CuSO_4 solution is taken as electrolyte.

Cu-atoms at anode lose electrons, come into the solution as Cu^{2+} ion. They migrate to cathode where they gain electrons and are deposited at cathode. In this way Cu atoms are migrated from anode to cathode. Impurities are left at the bottom of anode and pure Cu is obtained as thick cathode.



- (h) SHE has electrode potential equal to zero. When it is coupled with Zn, Zn has low reduction potential and has 0.76 V more tendency than hydrogen to show oxidation and act as anode. Hydrogen acts as cathode.

When Cu is coupled with hydrogen, Cu has high reduction potential and has 0.34 V more tendency than hydrogen to show reduction and acts as cathode. Then hydrogen acts as anode.



DO YOU KNOW?

1. The conversion of electrical energy into chemical energy in the electrolytic cell and the conversion of chemical energy into the electrical energy in the galvanic or voltaic cell is called electrochemistry.
2. The substance in which electricity can pass is called conductor.
3. The flow of charge in an electronic conductor is due to the flow of electrons.
4. The flow of charge in the electrolytic conductor is due to the movement of ions.
5. Oxidation state of O in OF_2 is +2.
6. Oxidation state of O in peroxides (Na_2O_2 , H_2O_2) is -1.
7. Oxidation state of O in super oxides (KO_2) is $-\frac{1}{2}$.
8. Oxidation is due to loss of electrons.
9. Reduction is due to the gain of electrons.
10. Oxidation takes place at the anode.
11. Reduction takes place at the cathode.
12. An electrochemical cell in which electric current is used to drive a non-spontaneous redox reaction is called electrolytic cell. For example: Down's cell, Nelson's cell.
13. An electrochemical cell in which electric current is produced as a result of spontaneous redox reaction is called galvanic cell or voltaic cell. For example: Lead storage battery dry cell, Daniel cell (Zn – Cu cell).
14. Lead storage battery has usually six cell, the voltage of each cell is 2 and the total voltage is 12 V.
15. 30% H_2SO_4 having density of 1.25 g cm^{-3} is used as electrolyte in the lead storage battery.
16. The electrolyte in fuel cell is KOH.
17. In the NICAD battery, anode is composed of cadmium.
18. In the silver oxide battery, cathode is Ag_2O and anode is made of zinc.
19. The voltage of alkaline battery is 1.5.
20. The oxidation and reduction potential of SHE is arbitrarily taken as zero.
21. In electrolytic cell, cathode is negatively charged and anode is positively charged.
22. In Galvanic cell, cathode is positively charged and anode is negatively charged.

