EXERCISE

MULTIPLE CHOICE QUESTIONS

1. Choose the correct answer.

i. The electrode through which the electron enters the electrolytic solution is

a) Anode b) Cathode c) Salt Bridge d) Electrolyte

ii. Oxidation number of S in $Na_2S_2O_3$ is

a) +1 b) +2 c) +3 d) +4

iii. In the electrolysis of molten $ZnCl_2$, the cathode reaction is $Zn^{+2} + 2e^- \rightarrow Zn$ what quantity of electricity is used for the production of half mole of Zn

a) 2 Coulombs b) 1F c) 2F d) 2 ampere

iv. How many moles of Cr will be produced by 1.5 faradays of electricity by the following reaction $Cr^{+3} + 3e^- \rightarrow Cr$

a) 0.1 b) 0.2 c) 0.03 d) 1.5 e) 0.5

v. $E°_{Sn} = -0.14V, E°_{Pb} = -0.13V$

a) Sn^{+2} can oxidize Pb b) Pb^{+2} can oxidize Sn

c) both can oxidize each other d) both can reduce each other

vi. A fuel call operates at _____ temperature

a) low b) medium c) room d) high

vii. The oxidation number of Cl in $HClO_4$ is

a) - 1 b) + 1 c) + 5 d) + 7

viii. In which of the following compounds oxidation number of N is + 5

a) NO_2 b) N_2O_4 c) N_2O_3 d) N_2O_5

ix. The passage of current through an electrolyte is due to the movement of

a) electrons b) anions c) cations d) ions

x. Oxidation number of an element in free state is

a) Negative b) positive c) zero d) ± 1

xi. Which of these is not true of an electrolyte.

a) It can conduct electricity in molten state

b) It can conduct electricity in the form of aqueous solution.

c) It can conduct electricity in the solid form.

d) It can be an acid, a base or a salt.

xii. In an electrolytic cell, the name of electrode which has more negative potential begins with the letter

a) A to B b) C to D c) K to O d) U to Z

xiii. Corrosion is an electrochemical process which requires

a) Oxygen b) Water c) Acidic vapours d) Basic vapours

1. a
2. a, b
3. a, b, c
4. a, b, c, d

xiv. A fuel cell is based upon the reaction between

a) Oxygen b) Gaseous fuel c) KOH d) Pt

1. a
2. a, b
3. a, b, c
4. a, b, c, d

Answers

i. b	ii. b	iii. b	iv. e	v. b	vi. d	vii. b
viii. d	ix. d	x. c	xi. c	xii. b	xiii. 3	xiv. 2

2. Explain, why?

(a) Reduction of 1 mole of each Zn^{+2} and Ag^{+} require different Faradays of electricity.

(b) It is not possible to measure the potential of an isolated half-cell.

(c) The life of a dry cell is shorter than that of an alkaline dry cell.

Ans: (a) One faraday is equal to the charge carried one mole of electron and it is equal to 96487 C.

One faraday is 1 mole of electrons.

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

We can see that this requires two moles of electrons per mole of zinc so 2 faradays or about 193000 C charge of electricity is required.

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

We can see that this requires one mole of electrons per mole of silver so 1 faraday or about

96500 C charge of electricity is required.

(b) A redox reaction always consists of oxidation and reduction. In the cell, the electrode potential develops due to oxidation at anode and reduction, at cathode. Therefore, it is not possible to measure the potential of an isolated half cell.

Example:

If only a zinc half-cell were constructed, no complete redox reaction can occur and so no electrical potential can be measured. It is only measured when another half-cell is combined with the zinc half-cell.

(c) The life of a dry cell is shorter than that of an alkaline dry cell because in alkaline dry cell, a moist past of KOH is used as electrolyte. That is why the alkaline dry cell lasts longer because the zinc anode corrodes less rapidly in basic conditions.

3. Write the spontaneous reaction for the following sets of half-reactions.

$$Au^{+3}(aq) + 3e^{-} \rightarrow Au(s) \qquad E^{\circ} = 1.5V$$

$$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l) \qquad E° = 0.096\,V$$

Solution:

(i) $Au^{+3}(aq) + 3e^- \rightarrow Au\,(s) \qquad E° = 1.5V$

(ii) $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l) E° = 0.096\,V$

The above equations shows that the reduction potential of first reaction is greater than the second reaction therefore the reduction will occur in the first reaction and oxidation will occur in the second reaction.

Cell reaction:

Cell reaction will occur if we reverse the second reaction and then add it into the first reaction.

At Cathode (Reduction):

$$Au^{+3}(aq) + 3e^- \rightarrow Au\,(s) \qquad E° = 1.5V$$

At Anode (Oxidation):

$$NO(g) + 2H_2O(l) \rightarrow NO_3^-(aq) + 4H^+(aq) + 3e^- E° = 0.096\,V$$

Net reaction: $Au^{+3}(aq) + NO(g) + 2H_2O(l) \rightarrow NO_3^-(aq) + 4H^+(aq) + Au\,(s)$

Cell potential:

$E° = E°_{cathode} - E°_{anode}$

$E° = (1.5\,V) - (0.96\,V)$

$E°_{.} = (1\,5\,V) - (0\,96\,V)$

$E°_{.} = 1\,596\,V$

As the cell potential is positive, therefore the reaction is spontaneous.

4. Explain the following with reasons.

(a) The oxidation potential of Zn is +0.76V and its reduction potential is -0.76V

(b) A salt bridge maintains the electrical neutrality in the cell.

(c) Na and K can display hydrogen from acids but Cu and Pt cannot.

(d) Lead storage battery is rechargeable battery.

(e) Zn plating saves Fe from corrosion.

Ans: **(a)** Oxidation potential of zinc metal is 0.76 V it shows that zinc can lose electron easily and oxidation reaction is feasible. The reduction potential of zinc metal is -0.76 V it shows that zinc cannot reduce easily.

(b) A Salt Bridge maintains the electrical neutrality in the cell because it only allows the movement of ions from one solution to the other without mixing the two solutions.

(c) It is observed that metal like Na, K can displace H_2 but metals like Cu, Pt cannot displace H_2.

Reason:

i. In active series Na and K are ranked above H_2 therefore they can displace hydrogen.

ii. In active series Cu and Pt are ranked below H_2 therefore they cannot displace hydrogen.

iii. The reduction potential values of Na and K are -2.714V and -2.925V respectively. That is why they can oxidized easily. Hence on reaction with an Acid, they displace the hydrogen easily.

The reduction potential of Cu is +0.521 V and that of Pt is +1.55V. Due to high values of reduction potentials. Cu is less reactive towards acids and Pt is nearly inert towards acids.

(d) During discharging of Lead Storage battery, H_2SO_4 is used up. Its density decreases. When both the electrodes are completely covered with $PbSO_4$ the battery ceases to deliver current.

Recharging the Battery:

i. Battery can be recharged by connecting its anode to the negative terminal of direct current and the cathode to the positive terminal of the direct current.

ii. Reverse chemical reactions occur at anode and cathode of the battery.

iii. Thus deposition of Pb on anode and PbO_2 on cathode takes place. In this way we can recharge the battery.

Redox reactions during recharging:

The reactions of recharging of battery are as follows:

At cathode:

$$PbSO_{4(s)} + 2e^- \rightarrow Pb_{(s)} + SO_4^{-2}{}_{(aq)}$$

At anode: $$\mathbf{PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow PbO_{2(s)} + SO_4^{-2}{}_{(aq)} + 4H^+{}_{(aq)} + 2e^-}$$

Net reaction: $$\mathbf{2PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow Pb_{(s)} + PbO_{2(s)} + 4H^+{}_{(aq)} + 2SO_4^{-2}{}_{(aq)}}$$

After recharging H_2SO_4 solution is concentrated again bringing density to its initial value of 1.25 g cm^{-3}.

5. What kind of reaction occurs at the negative terminal of the galvanic cell?

Ans: Oxidation reaction will occur at the negative terminal of the galvanic cell.

Explanation:

In Zn-Cu galvanic cell Zn has greater tendency to lose electrons than Cu. Therefore, Zn electrode acquires negative charge relative to Cu electrode. The

electrons flow from Zn electrode through the external circuit to Cu electrode. Thus, zinc will act as the negative terminal.

Hence at negative terminal (anode) of cell an oxidation reaction will occur.

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

6. Corrosion is often accelerated where the coating on the body of a car has begun to crack, Explain.

Ans: Body of a car is made up of iron and it is coated with paint. Paint coating save the body of the car from rust. Corrosion is often accelerated where the coating on the body of a car has begun to crack it is because the metal of the car comes in contact with oxygen and moisture and other harmful agents and following reaction would take place.

Reduction half reaction: $O_{2(g)} + H^+_{(aq)} + e^- \rightarrow 2H_2O_{(l)}$

Oxidation half reaction: $Fe \rightarrow Fe^{+2} + e^-$ **(oxidation reaction)**

Conclusion:

Hence oxygen oxidizes Fe to Fe^{+2} which are further oxidized to Fe^{+3} the Fe^{+3} form various insoluble hydrated oxides known as rust and thus the corrosion will be accelerated.

7. How many hours would electroplating have to be continued at the rate of 5amperes if 75g of copper is to be deposited from $CuSO_4$ solution?

Solution:

Mass of Cu to be produced = 75 g

Atomic mass of Cu = 63.5 $g\ mole^{-1}$

$$Number\ of\ moles\ of\ Cu = \frac{75g}{63.5g\ mole^-} = 1.1811\ moles$$

Reaction at cathode:

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

The above reaction indicate that for every 2 moles of electrons, 1 mole of Cu will produce.

2 F = 2 moles of electrons = 1 mole of Cu.

1 mole of Cu is produced by charge = 2 F

1.1811 moles of Cu is produced by charge = 2 × 1.1811

= 2.3622 F

As we know that 1 F = 96500 C

2.3622 F = 96500 C × 2.3622 = 227952.3 C

Time = ?

Charge = 227952.3 C

Current = 5 A

$$Current = \frac{charge}{time}$$

$$Time = \frac{charge}{current} = \frac{227952.3}{5} = 45590.46\ s$$

$$Time = \frac{45590.46}{3600} = 12.664\ hours$$

8. What is the difference between the following:

(a) A galvanic and electrolytic cell

(b) Oxidation half-reaction and reduction half reaction

Ans: (a) A galvanic and electrolytic cell:

Galvanic Cell	Electrolytic cell

i. A Galvanic cell converts chemical energy into electrical energy	i. An electrolytic cell converts electrical energy into chemical energy
ii. Here, the Redox reaction is spontaneous and is responsible for the production of electrical energy.	ii. The Redox reaction is not spontaneous and electrical energy has to be supplied to initiate the reaction.
iii. The two half-cells are setup in different containers, being connected through the Salt Bridge or porous partition.	iii. Both the electrodes are placed in a same container in the solution of molten electrolyte.
iv. Here the anode is negative and Cathode is positive electrode. The reaction at anode is oxidation and that at the cathode is reduction.	iv. Here the anode is positive and Cathode is negative electrode. The reaction at anode is oxidation and that at the cathode is reduction.
v. The electrons are supplied by the species getting oxidized. They move from anode to cathode in the external circuit.	v. The external battery supplies the electrons. They enter through the cathode and come out through the anode.

(b) Oxidation half-reaction and reduction half reaction:

Oxidation half-reaction	Reduction half reaction

i. Reaction taking place in oxidation half-cell is called oxidation half reaction.	i. The reaction taking place in reduction half-cell is called reduction half reaction.
ii. In oxidation half-reaction, the species is oxidized by losing electrons.	ii. In reduction half-reaction, the species is reduced by gaining electrons.
iii. This process occurs at anode	iii. This process occurs at cathode
iv. Oxidation half-reaction $Zn_{(s)} \rightarrow Zn^{+2}{}_{(aq)} + 2e^-$	iv. Reduction half reaction $Cu^{+2}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

9. Sketch the Galvanic cell based on the following overall reaction, show direction of electron flow and identity the anode, write balanced reaction. Assume all concentrations are 1.0 M

(a) $Cr^{+2}(aq) + CI_2(aq) \rightarrow Cr_2O_7{}^{2}(aq) + CI^-(aq)$

(b) $Cu^{+2}(aq) + Mg(s) \rightarrow Mg^{+2}(aq) + Cu(s)$

Ans: (a) $Cr^{+2}(aq) + CI_2(aq) \rightarrow Cr_2O_7{}^{2}(aq) + CI^-(aq)$

i. In this reaction Cl atom changes its Oxidation State from zero in Cl_2 to -1 in CI^-. It is a reduction process as it gain 1e per Cl atom.

ii. In this reaction Cr atom changes its Oxidation State from +3 in Cr^{+3} to +6 in $Cr_2O_7{}^{2}$. It is a oxidation process as it loss 3e per Cr atom.

At anode: (oxidation half reaction)

$$2Cr^{+2}(aq) + 7H_2O \rightarrow Cr_2O_7{}^{2}(aq) + 14H^+ + 6e^- \quad \text{.....(i)}$$

At cathode: (reduction half reaction)

$$C \ldots + 2e \rightarrow 2CI^-(aq) \quad \text{.....(ii)}$$

Electron flow:

Thus the electron flows from anode to cathode.

Balancing:

To balance loss and gain of electrons multiply Eq. (ii) by 3 and then add to Eq (i) to get overall cell reaction.

Oxidation half reaction: $2Cr^{+3}(aq) + 7H_2O \rightarrow Cr_2O_7^{-2}(aq) + 14H^+ + 6e^-$

Reduction half reaction: $3Cl_{2(aq)} + 6e^- \rightarrow 6Cl^{-1}_{(aq)}$

Overall cell reaction:

$$2Cr^{+3}(aq) + 7H_2O + 3Cl_{2(aq)} \rightarrow Cr_2O_7^{-2}(aq) + 14H^+ + 6Cl^{-1}_{(aq)}$$

(b) $Cu^{+2}(aq) + Mg(s) \rightarrow Mg^{+2}(aq) + Cu(s)$

i. The Cu atom changes its Oxidation State from +2 in Cu^{+2} to 0 in Cu. It is a reduction process as it gain 2e per atom.

ii. In this reaction Mg atom changes its Oxidation State from 0 in Mg to +2 in Mg^{+2}. It is a oxidation process as it gain2e per atom.

At anode: (oxidation half reaction)

$$Mg(s) \rightarrow Mg^{+2} + 2e^- \qquad \text{.....(i)}$$

At cathode: (reduction half reaction)

$$Cu_2(aq) + 2e^- \rightarrow Cu_{(aq)} \qquad \text{.....(ii)}$$

Electron flow:

Thus the electron flows from anode to cathode.

Balancing:

Write equation (i) and (ii) as it is

Oxidation half reaction: $Mg(s) \rightarrow Mg^{+2} + 2e^-$

Reduction half reaction: $Cu_2(aq) + 2e^- \rightarrow Cu_{(aq)}$

Net reaction: $Cu^{+2}(aq) + Mg(s) \rightarrow Mg^{+2} + Cu_{(aq)}$

Q.10: Calculate E° for the following cells, which reactions are spontaneous as written under standard conditions.

Ans: (a) $2Ag^+(aq) + Cu(s) \rightarrow Cu^{+2}(aq) + 2Ag(s)$

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

(c) $Al(s) + 3H^+(aq) \rightarrow Al^{+3}(aq) + \frac{3}{2}H_2(g)$

Solution:(a) $2Ag^+(aq) + Cu(s) \rightarrow Cu^{+2}(aq) + 2Ag(s)$

Reduction half reaction:

Ag atom changes its oxidation state from +1 in Ag^+ to zero in Ag. It is because of the gain of $1e^-$. Thus it is a reduction process.

At cathode: (Reduction half reaction)

$$2Ag^+(aq) + 2e^- \rightarrow 2Ag(aq)$$

Oxidation half reaction:

Cu atom changes its oxidation state from zero in Cu to +2 in Cu^{+2}. It is because of the loss of $2e^-$. Thus it is an oxidation process.

At anode: (Oxidation half reaction)

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

Cell potential:

The standard reduction potentials of Cu and Ag are

$$E^\circ_{Cu} = +0.34V \text{ and } E^\circ_{Ag} = +0.80\,V$$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cell} = E^\circ_{Ag} - E^\circ_{Cu}$$

$$E^\circ_{cell} = (0.80\ V) - (0.34\ V)$$

$$E^\circ_{cell} = 0.46\ V$$

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

Reduction half reaction:

Zn atom changes its oxidation state from +2 in Zn^{+2} to zero in Zn. It is because of the gain of $2e^-$. Therefore it is a reduction process.

At cathode: (Reduction half reaction)

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

Oxidation half reaction:

Ni atom changes its oxidation state from zero in Ni to +2 in Ni^{+2}. It is because of the loss of $2e^-$. Therefore it is an oxidation process.

At anode: (Oxidation half reaction)

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

Cell potential:

The standard reduction potentials of Ni and Zn are

$$E^\circ_{Ni} = -0.25\ V \text{ and } E^\circ_{Zn} = -0.76\ V$$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

$$E^\circ_{cell} = E^\circ_{Zn} - E^\circ_{Ni}$$

$$E^\circ_{cell} = (-0.76) - (-0.25)$$

$$E^\circ_{cell} = -0.51$$

As the cell potential is negative. Thus this reaction is not spontaneous (not feasible)

(c) $Al(s) + 3H^+(aq) \rightarrow Al^{+3}(aq) + \frac{3}{2}H_2(g)$

Oxidation half reaction:

In this reaction, the Al atom changes its oxidation state from zero in Al to +3 in Al^{+3}. It is because of the loss of $3e^-$. Therefore it is an oxidation process.

Anode: (Oxidation half reaction)

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

Reduction half reaction:

The H atom changes its oxidation state from +1 in H^+ to zero in H_2. This is due to gain of $1e^-$. Thus it is a reduction process.

At cathode: (Reduction half reaction)

$$3H^+(aq) + 3e^- \rightarrow \frac{3}{2}H_2(g)$$

Cell potential:

Standard reduction potentials of hydrogen and Al are

$$E°_{Al} = -1.67\ V \text{ and } E°_{H} = 0.00\ V$$

$$E°_{cell} = E°_{cathode} - E°_{anode}$$

$$E°_{cell} = E°_{H} - E°_{Al}$$

$$E°_{cell} = (0.00) - (-1.67)$$

$$E°_{cell} = +1.67\ V$$

As the cell potential is positive, Thus, this reaction is spontaneous (feasible)

Q.11. An electroplating apparatus is used to coat jewelry with gold. What mass of gold can be deposited from a solution that contains $\{Au(Cn)_4\}^{-1}$ ion if a current of 5.0amperes flows for 30 minutes, following half-reaction occurs.

$$\{Au(Cn)_4\}^{-1}(aq) + 3e^- \rightarrow Au(s) + 4CN^-(aq)$$

Solution: Charge calculation:

Time $= 30\ minutes = 30 \times 60 = 1800s$

Current $= 5\ A$

As

$$Charge = \frac{charge}{Time(s)}$$

$$Charge = current \times Time(s) = 5\ A \times 1800s = 9000\ C$$

As

96500 C $= 1F$

9000 C $= \frac{1F}{96500\ C} \times 9000\ C = 9.326 \times 10^{-2}\ F$

Reaction at cathode:

$$\{Au(Cn)_4\}^{-1}(aq) + 3e^- \rightarrow Au(s) + 4CN^-(aq)$$

It indicates that every 3 moles of electrons, 1 mole of Au will produce.

1 F = 1 mole of electron.

3 F = 3 moles of electrons = 1 mole of Au.

So,

3 F charge produces Au = 1 mole

9.326×10^{-2} F charge produces $= \frac{1}{3} \times 9.326 \times 10^{-2}$

$$= 3.109 \times 10^{-2}\ moles$$

Atomic mass of Au $= 179\ g\ mole^{-1}$

1 mole of Au $= 179\ g$

$3.109 \times 10^{-2}\ moles = 179 \times 3.109 \times 10^{-2} = 5.565\ g$

Q.12. Determine oxidation number of all the atoms in the following

(a) K_2CO_3 (b) P_4 (c) XeF_4 (d) F_2O

(e) PCI_5 (f) NH_3 (g) Fe_3O_4

Ans:

(a) K_2CO_3

Oxidation number of K $= +1$

Oxidation number of O $= -2$

Oxidation number of C $= x$

$$K_2CO_3$$

$$2(+1) + x + 3(-2) = 0$$

$$x - 4 = 0$$

$$x = +4$$

(b) P_4

In free state Oxidation state of an element is zero. Thus oxidation state of P_4 is zero.

(c) XeF_4

Oxidation number of F $= -1$

Oxidation number of Xe $= x$

$$XeF_4$$

$$x + 4(-1) = 0$$

$$x - 4 = 0$$

$$x = +4$$

(d) F_2O

Oxidation number of F $= -1$

Oxidation number of O $= x$

$$F_2O$$

$$2(-1) + x = 0$$

$$-2 + x = 0$$

$$x = +2$$

(e) $\mathbf{PCI_5}$

Oxidation number of CI $= -1$

Oxidation number of P $= x$

$$PCI_5$$

$$x + 5(-1) = 0$$

$$x - 5 = 0$$

$$x = +5$$

(f) NH_3

Oxidation number of H $= +1$

Oxidation number of N $= x$

$$NH_3$$

$$x + 3(+1) = 0$$

$$x + 3 = 0$$

$$x = -3$$

(g) Fe_3O_4

Oxidation number of O $= -2$

Oxidation number of Fe $= x$

$$Fe_3O_4$$

$$3x + 4(-2) = 0$$

$$3x - 8 = 0$$

$$3x = +8$$

$$x = +\frac{8}{3}$$

Q.13: Balance the following half-reactions

(a) $Cu^{+} \rightarrow CuO$ **(b)** $I \rightarrow IO_3^{-1}$

(c) $S \rightarrow H_2S$ **(d)** $SO_4^{-2} \rightarrow SO_2$

Ans: (a) $Cu^{+} \rightarrow CuO$

(i) Balancing of oxygen:

There is one O atom on the right and none on the left. So we will add one H_2O on the L.H.S. to balance O atoms.

$$Cu^{+} + H_2O \rightarrow CuO$$

(ii) Balancing of hydrogen:

There is 2H atoms on the left and none on the right. So we will add $2H^{+}$ on the R.H.S. to balance H.

$$Cu^{+} + H_2O \rightarrow CuO + 2H^{+}$$

(iii) Balance Charge:

Now balance charges the left side has mono-positive charge corresponding to +1. The right side has 2 mono-positive charges corresponding to +2. Thus, right side needs $1e^-$

$$Cu^+ + H_2O \rightarrow CuO + 2H^+ + 1e^-$$

(b) $I \rightarrow IO_3^{-1}$

(i) Balancing of oxygen:

There are 3 O atom on the right and none on the left. So we will add $3H_2O$ on the L.H.S. to balance O atoms.

$$I + 3H_2O \rightarrow IO_3^{-1}$$

(ii) Balancing of hydrogen:

There is 6H atoms on the left and none on the right. So we will add $6H^+$ on the R.H.S. to balance H.

$$I + 3H_2O \rightarrow IO_3^{-1} + 6H^+$$

(iii) Balance Charge:

Now balance charges the left side has no charge corresponding to 0. The right side has 6 mono-positive and mono-negative charge corresponding to -1+6(+1)=+5. Thus, right side needs $5e^-$

$$I + 3H_2O \rightarrow IO_3^{-1} + 6H^+ + 5e^-$$

(c) $S \rightarrow H_2S$

(i) There are 2 hydrogen atoms on the right side and none on the left side. So add $2H^+$ on the L.H.S. to balance hydrogen.

$$S + 2H^+ \rightarrow H_2S$$

(iii) Balance Charges:

Now balance charges the left side has 2 mono-positive charge correspond to 2(+1)=+2. No, charges present on right hand side Thus, left hand side needs $2e^-$

$$S + 2H^+ + 2e^- \rightarrow H_2S$$

(d) $SO_4^{-2} \rightarrow SO_2$

(i) Balancing of oxygen:

There are 4 Oxygen atoms on the left and 2 oxygen atoms on the right to balance oxygen add $2H_2O$ on the right side.

$$SO_4^{-2} \rightarrow SO_2 + 2H_2O$$

(ii) Balancing of hydrogen:

There is 4H atoms on the right and none on the left to balance hydrogen add $2H^+$ on the L.H.S.

$$SO_4^{-2} + 4H^+ \rightarrow SO_2 + 2H_2O$$

(iii) Balance Charge:

There are 4 mono-positive and1 di-negative charge on the left side correspond to (-2) +4(+1) = +2. The right side has no charge Thus, left side $2e^-$

$$SO_4^{-2} + 4H^+ + 2e^- \rightarrow SO_2 + 2H_2O$$

Q.14: Balance the following reactions by ion electron method.

(a) $SO_2 + MnO_4^- \rightarrow Mn^{+2} + HSO_4^-$

(b) $BrO_3^{-2} + SO_2 \rightarrow HSO_4^- + Br^-$

(c) $2N_2O_4 \rightarrow NO_3^- + NO$

(d) $Sn^{+2} + Fe^{+3} \rightarrow Sn^{+4} + Fe^{+2}$

(e) $H_2O_2 + MnO_4^- \rightarrow Mn^{+2} + O_2$

(f) $Cr_2O_7^{-2} + CI^{-1} \rightarrow Cr^{+3} + CI_2$

(g) $IO_3^{-1} + AsO_3^{-3} \rightarrow I^{-1} + AsO_4^{-3}$

(h) $H_3AsO_3 + Cr_2O_7^{-2} \rightarrow H_3AsO_4 + Cr^{+3}$

(i) $Cu + NO_3^- \rightarrow Cu^{+2} + NO$

Solution:

(a) $SO_2 + MnO_4^- \rightarrow Mn^{+2} + HSO_4^-$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$MnO_4^{-1} \rightarrow Mn^{+2}$

Oxidation half reaction:

$SO_2 \rightarrow HSO_4^{-1}$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) **Balance Mass:**

Mn atoms are already balanced.

$MnO_4^{-1} \rightarrow Mn^{+2}$

Add $4H_2O$ on right side to balance O atoms.

$MnO_4^{-1} \rightarrow Mn^{+2} + 4H_2O$

Add $8H^+$ on L.H.S. to balance H.

$MnO_4^{-1} + 8H^+ \rightarrow Mn^{+2} + 4H_2O$

(ii) **Balance Charge:**

The total charge on L.H.S. is -1+8(+1) = +7. The total charge on R.H.S. is +2. Thus, add $+5e^-$ on the L.H.S.

$MnO_4^{-1} + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$ .....(i)

Oxidation half reaction:

(i) **Balance Mass:**

S atoms are already balanced.

$$SO_2 \rightarrow HSO_4^{-1}$$

Add $4H_2O$ on left hand side to balance O atoms.

$$SO_2 + 4H_2O \rightarrow HSO_4^{-1}$$

Add $3H^+$ on R.H.S. to balance H atoms.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is (-1)+3(+1)=+2. Thus, add $2e^-$ on the R.H.S.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $5e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (1) by 2. Eq (2) by 5 and then add the resulting equations.

$$2MnO_4^{-1} + 16H^+ + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5SO_2 + 10H_2O \rightarrow 5HSO_4^{-1} + 15H^+ + 10e^-$$

$$2MnO_4^{-1} + 16H^+ + 10e^- + 5SO_2 + 10H_2O \rightarrow 2Mn^{+2} + 15HSO_4^{-1} + 15H^+ + 10e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $10e^-$, $15H^+$ and $8H_2O$. Cancel this out from both the sides.

$$2MnO_4^{-1} + H^+ + 5SO_2 + 2H_2O \rightarrow 2Mn^{+2} + 5HSO_4^{-1}$$

(b) $BrO_3^{-2} + SO_2 \rightarrow HSO_4^- + Br^-$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$BrO_3^{-2} \rightarrow Br^{-1}$$

Oxidation half reaction:

$$SO_2 \rightarrow HSO_4^{-1}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Br atoms are already balanced.

$$BrO_3^{-2} \rightarrow Br^{-1}$$

Add $3H_2O$ on right side to balance O atoms.

$$BrO_3^{-2} \rightarrow Br^{-1} + 3H_2O$$

Add $6H^+$ on L.H.S. to balance H.

$$BrO_3^{-2} + 6H^+ \rightarrow Br^{-1} + 3H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -2+6(+1) = +4. The total charge on R.H.S. is -1. Thus, add $+5e^-$ on the L.H.S.

$$BrO_3^{-2} + 6H^+ + 5e^- \rightarrow Br^{-1} + 3H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

S atoms are already balanced.

$$SO_2 \rightarrow HSO_4^{-1}$$

Add $3H_2O$ on left hand side to balance O atoms.

$$SO_2 + 4H_2O \rightarrow HSO_4^{-1}$$

Add $3H^+$ on R.H.S. to balance H atoms.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is (-1)+3(+1)=+2. Thus, add $2e^-$ on the R.H.S.

$$SO_2 + 2H_2O \rightarrow HSO_4^{-1} + 3H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $5e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (1) by 2. Eq (2) by 5 and then add the resulting equations.

$$2BrO_3^{-2} + 12H^+ + 10e^- \rightarrow 2Br^{-1} + 6H_2O$$

$$5SO_2 + 10H_2O \rightarrow 5HSO_4^{-1} + 15H^+ + 10e^-$$

$$2BrO_3^{-2} + 12H^+ + 10e^- + 5SO_2 + 10H_2O \rightarrow 2Br^{-1} + 6H_2O + 5HSO_4^{-1} + 15H^+ + 10e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $10e^-$, $12H^+$ and $6H_2O$. Cancel this out from both the sides.

$$2BrO_3^{-2} + 5SO_2 + 4H_2O \rightarrow 2Br^{-1} + 5HSO_4^{-1} + 3H^+$$

(c) $2N_2O_4 \rightarrow NO_3^- + NO$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$N_2O_4 \rightarrow NO$$

Oxidation half reaction:

$$N_2O_4 \rightarrow NO_3^{-1}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before NO on R.H.S. to balance N atoms.

$$N_2O_4 \rightarrow 2NO$$

Add $2H_2O$ on right side to balance O atoms.

$$N_2O_4 \rightarrow 2NO + 2H_2O$$

Add $4H^+$ on L.H.S. to balance H.

$$N_2O_4 + 4H^+ \rightarrow 2NO + 2H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is 4(+1) = +4. The R.H.S. has no charge. Thus, add $4e^-$ on the L.H.S.

$$N_2O_4 + 4H^+ + 4e^- \rightarrow 2NO + 2H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

Use 2 before NO_3^{-1} on R.H.S. to balance N atoms.

$$N_2O_4 \rightarrow 2NO_3^{-1}$$

Add $2H_2O$ on left hand side to balance O atoms.

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1}$$

Add $4H^+$ on R.H.S. to balance H atoms.

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1} + 4H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is 2(-1)+4(+1)=+2. Thus, add $2e^-$ on the R.H.S.

$$N_2O_4 + 2H_2O \rightarrow 2NO_3^{-1} + 4H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $4e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (2) by 2 and then add to Eq. (1)

$$N_2O_4 + 4H^+ + 4e^- \rightarrow 2NO + 2H_2O$$

$$2N_2O_4 + 4H_2O \rightarrow 4NO_3^{-1} + 8H^+ + 4e^-$$

$$3N_2O_4 + 4H^+ + 4H_2O + 4e^- \rightarrow 2NO + 2H_2O + 4NO_3^{-1} + 8H^+ + 4e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $4e^-$, $4H^+$ and $2H_2O$. Cancel this out from both the sides.

$$3N_2O_4 + 2H_2O \rightarrow 2NO + 4NO_3^{-1} + 4H^+$$

(d) $Sn^{+2} + Fe^{+3} \rightarrow Sn^{+4} + Fe^{+2}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

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

Oxidation half reaction:

$$Sn^{+2} \rightarrow Sn^{+4}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Fe atoms are already balanced.

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

(ii) Balance Charge:

The total charge on L.H.S. is +3. The total charge on R.H.S. is +2. Thus, add $4e^-$ on the L.H.S.

$$Fe^{+3} + 1e^- \rightarrow Fe^{+2} \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

Sn atoms are already balanced.

$$Sn^{+2} \rightarrow Sn^{+4}$$

(ii) Balance Charge:

The total charge on L.H.S. is +2. The total charge on R.H.S. is +4. Thus, add $2e^-$ on the R.H.S.

$$Sn^{+2} \rightarrow Sn^{+4} + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $1e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (1) by 2 and then add to Eq. (2)

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

$$Sn^{+2} \rightarrow Sn^{+4} + 2e^-$$

$$2Fe^{+3} + 2e^- + Sn^{+2} \rightarrow 2Fe^{+2} + Sn^{+4} + 2e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $4e^-$. $4H^+$ and $2H_2O$. Cancel this out from both the sides.

$$2Fe^{+3} + Sn^{+2} \rightarrow 2Fe^{+2} + Sn^{+4}$$

(e) $H_2O_2 + MnO_4^- \rightarrow Mn^{+2} + O_2$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

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

Oxidation half reaction:

$$H_2O_2 \rightarrow O_2$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Mn atoms are already balanced.

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

Add $4H_2O$ on right side to balance O atoms.

$$MnO_4^- \rightarrow Mn^{+2} + 4H_2O$$

Add $8H^+$ on L.H.S. to balance H.

$$MnO_4^- + 8H^+ \rightarrow Mn^{+2} + 4H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+8(+1) = +7. The total charge on R.H.S. is +2. Thus, add $5e^-$ on the L.H.S.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

O atoms are already balanced.

$$H_2O_2 \rightarrow O_2$$

Add $2H^+$ on R.H.S. to balance H atoms.

$$H_2O_2 \rightarrow O_2 + 2H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is 2(+1)=+2. Thus, add $2e^-$ on the R.H.S.

$$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $5e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (1) by 2 by 5and then add the resulting equations.

$$2MnO_4^{-1} + 16H^+ + 10e^- \rightarrow 2Mn^{+2} + 8H_2O$$

$$5H_2O_2 \rightarrow 5O_2 + 10H^+ + 10e^-$$

$$2MnO_4^{-1} + 16H^+ + 10e^- + 5H_2O_2 \rightarrow 2Mn^{+2} + 8H_2O + 5O_2 + 10H^+ + 10e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $10e^-$ and $10H^+$. Cancel this out from both the sides.

$$2MnO_4^{-1} + 6H^+ + 5H_2O_2 \rightarrow 2Mn^{+2} + 8H_2O + 5O_2$$

(f) $Cr_2O_7^{-2} + Cl^{-1} \rightarrow Cr^{+3} + Cl_2$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$Cr_2O_7^{-2} \rightarrow Cr^{+3}$$

Oxidation half reaction:

$$Cl^{-1} \rightarrow Cl_2$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before Cr^{+3} on R.H.S. to balance Cr atoms.

$$Cr_2O_7^{-2} \rightarrow 2Cr^{+3}$$

Add $7H_2O$ on right side to balance O atoms.

$$Cr_2O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$$

Add $14H^+$ on L.H.S. to balance H atoms.

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -2+14(+1) = +12. The total charge on R.H.S. is 2(+3) = +6. Thus, add $6e^-$ on the L.H.S.

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

Use 2 before Cl^{-1} ions on R.H.S. to balance Cl atoms.

$$2Cl^{-1} \rightarrow Cl_2$$

(ii) Balance Charge:

The total charge on L.H.S. is 2(-1)= -2. The R.H.S. has no charge Thus, add $2e^-$ on the R.H.S.

$$2Cl^{-1} \rightarrow Cl_2 + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $6e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (2) by 3 and then add to Eq. (1)

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$$

$$6Cl^{-1} \rightarrow 3Cl_2 + 2e^-$$

$$Cr_2O_7^{-2} + 14H^+ + 6Cl^{-1} + 6e^- \rightarrow 2Cr^{+3} + 7H_2O + 3Cl_2 + 2e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $6e^-$.Cancel this out from both the sides.

$$Cr_2O_7^{-2} + 14H^+ + 3Cl_2 + 6Cl^{-1} \rightarrow 2Cr^{+3} + 7H_2O + 3Cl_2$$

(g) $IO_3^{-1} + AsO_3^{-3} \rightarrow I^{-1} + AsO_4^{-3}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$IO_3^{-1} \rightarrow I^{-1}$$

Oxidation half reaction:

$$AsO_3^{-3} \rightarrow AsO_4^{-3}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

I atoms are already balanced.

$$IO_3^{-1} \rightarrow I^{-1}$$

Add $3H_2O$ on right side to balance O atoms.

$$IO_3^{-1} \rightarrow I^{-1} + 3H_2O$$

Add $6H^+$ on L.H.S. to balance H.

$$IO_3^{-1} + 6H^+ \rightarrow I^{-1} + 3H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+6(+1) = +5. The total charge on R.H.S. is -1. Thus, add $5e^-$ on the L.H.S.

$$IO_3^{-1} + 6H^+ + 6e^- \rightarrow I^{-1} + 3H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

S atoms are already balanced.

$$AsO_3^{-3} \rightarrow AsO_4^{-3}$$

Add H_2O on left side to balance O atoms.

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3}$$

Add $2H^+$ on L.H.S. to balance H atoms.

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3} + 2H^+$$

(ii) Balance Charge:

The total charge on L.H.S. is -3. The total charge on R.H.S. is -3+2(+1)= -1. Thus, add $2e^-$ on the R.H.S.

$$AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3} + 2H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $6e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (2) by 3 and then add to Eq (1).

$$IO_3^{-1} + 6H^+ + 6e^- \rightarrow I^{-1} + 3H_2O$$

$$3AsO_3^{-3} + 3H_2O \rightarrow 3AsO_4^{-3} + 6H^+ + 6e^-$$

$$IO_3^{-1} + 6H^+ + 6e^- + 3AsO_3^{-3} + 3H_2O \rightarrow I^{-1} + 3H_2O + 3AsO_4^{-3} + 6H^+ + 6e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $6e^-$, $6H^+$ and $3H_2O$. Cancel this out from both the sides.

$$IO_3^{-1} + 3AsO_3^{-3} \rightarrow I^{-1} + 3AsO_4^{-3}$$

(h) $H_3AsO_3 + Cr_2O_7^{-2} \rightarrow H_3AsO_4 + Cr^{+3}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$Cr_2O_7^{-2} \rightarrow Cr^{+3}$$

Oxidation half reaction:

$$H_3AsO_3 \rightarrow H_3AsO_4$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

Use 2 before Cr^{+3} on R.H.S. to balance Cr atoms.

$$Cr_2O_7^{-2} \rightarrow Cr^{+3}$$

Add $7H_2O$ on right side to balance O atoms.

$$Cr_2O_7^{-2} \rightarrow Cr^{+3} + 7H_2O$$

Add $14H^+$ on L.H.S. to balance H.

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -2+14(+1) = +12. The total charge on R.H.S. is 2(+3) = +6. Thus, add $6e^-$ on the L.H.S.

$$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

As atoms are already balanced.

$$H_3AsO_3 \rightarrow H_3AsO_4$$

Add $1\ H_2O$ on left side to balance O atoms.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4$$

Add $2H^+$ on R.H.S. to balance H atoms.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4 + 2H^+$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is +2. Thus, add $2e^-$ on the R.H.S.

$$H_3AsO_3 + H_2O \rightarrow H_3AsO_4 + 2H^+ + 2e^- \quad \text{.....(ii)}$$

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $6e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (2) by 3 and then add to Eq (1).

$$Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$$

$$3H_3AsO_3 + 3H_2O \rightarrow 3H_3AsO_4 + 6H^+ + 6e^-$$

$$Cr_2O_7^{-2} + 14H^+ + 3H_3AsO_3 + 3H_2O + 6e^- \rightarrow 2Cr^{+3} + 7H_2O + 3H_3AsO_4 + 6H^+ + 6e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $6e^-$, $6H^+$ and $3H_2O$. Cancel this out from both the sides.

$$Cr_2O_7^{-2} + 8H^+ + 3H_3AsO_3 \rightarrow 2Cr^{+3} + 4H_2O + 3H_3AsO_4$$

(i) $\boldsymbol{Cu + NO_3^- \rightarrow Cu^{+2} + NO}$

Step 1: Split the reaction into two half reactions.

Reduction half reaction:

$$NO_3^{-1} \rightarrow NO$$

Oxidation half reaction:

$$Cu \rightarrow Cu^{+2}$$

Step 2: Balance each half reaction separately.

Reduction half reaction:

(i) Balance Mass:

N atoms are already balanced.

$$NO_3^{-1} \rightarrow NO$$

Add $2H_2O$ on right side to balance O atoms.

$$NO_3^{-1} \rightarrow NO + 2H_2O$$

Add $6H^+$ on L.H.S. to balance H.

$$NO_3^{-1} + 4H^+ \rightarrow NO + 2H_2O$$

(ii) Balance Charge:

The total charge on L.H.S. is -1+4(+1) = +3. The R.H.S. has no charge. Thus, add $5e^-$ on the L.H.S.

$$NO_3^{-1} + 4H^+ + 3e^- \rightarrow NO + 2H_2O \quad \text{.....(i)}$$

Oxidation half reaction:

(i) Balance Mass:

Cu atoms are already balanced.

$$Cu \rightarrow Cu^{+2}$$

(ii) Balance Charge:

The L.H.S. has no charge. The total charge on R.H.S. is +2. Thus, add $2e^-$ on the R.H.S.

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

Step 3: Balance the loss and gain of the electrons in both half reactions and add them.

The reduction half reaction has $3e^-$ while the oxidation half-reaction $2e^-$. Thus multiply Eq (1) by 2, Eq (2) by 3 and then add the resulting equations.

$$2NO_3^{-1} + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O$$

$$3Cu \rightarrow Cu^{+2} + 6e^-$$

$$3Cu + 2NO_3^{-1} + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O + 3Cu^{+2} + 6e^-$$

Step 4: Strike out duplications of species on both the sides.

Duplications are $6e^-$. Cancel this out from both the sides.

$$2NO_3^{-1} + 8H^+ + 3Cu \rightarrow 2NO + 4H_2O + Cu^{+2}$$

15. Balance the following equations by oxidation number method.

(i) $MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$

(ii) $HNO_3 + HI \rightarrow NO + I_2 + H_2O$

(iii) $Ag + H_2S + O_2 \rightarrow Ag_2S + H_2O$

(iv) $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO + H_2O$

(v) $Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

(vi) $Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$

(vii) $HI + H_2SO_4 \rightarrow I_2 + SO_4 + H_2O$

(viii) $NaClO_2 + Cl_2 \rightarrow NaCl + ClO_2$

(ix) $Na + H_2O \rightarrow NaOH + H_2$

Solution:

(i) $MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$

Step 1: Assign oxidation number to all the atoms involved in the equation.

+4 +2(-2) +1 -1 +2 2(-1) +2 -2 0

$MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$

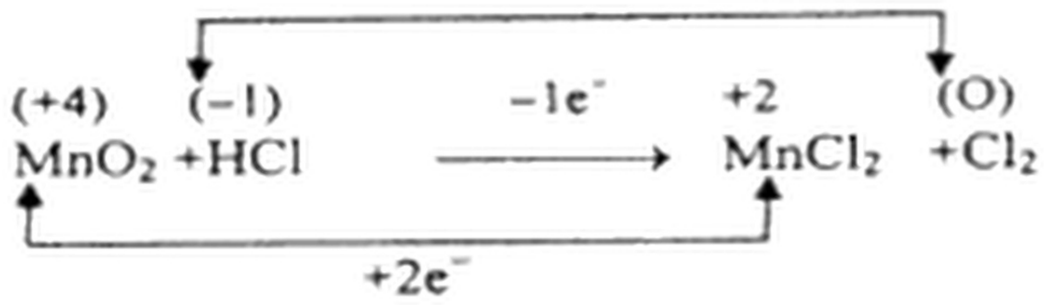
Step 2: Identify the elements undergoing a change in oxidation number

Some Cl atoms change their oxidation state from to -1 in HCl to zero in Cl_2. This is due to loss of + $1e^-$ per Cl atom. This is oxidation. However, some Cl atoms retain their oxidation state at -1 in HCl and $MnCl_2$. So, write HCl twice on L.H.S.

$HCl + MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$

The Mn atom changes its oxidation state from +4 in MnO_2 to +2 in $MnCl_2$. This is due to gain of $2e^-$ per Mn atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons, use multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $1e^-$ loss by 2 and $2e^-$ gain by 1.

Multiply the coefficient of HCl by 2 in order to equalize the number of Cl atoms for electrons loss.

$$HCl + MnO_2 + 2HCl \rightarrow MnCl_2 + H_2O + Cl_2$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

To balance Cl atoms multiply coefficient of other HCl by 2.

$$2HCl + MnO_2 + 2HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$$

Or $$MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$$

This is the balanced equation.

(ii) $\boldsymbol{HNO_3 + HI \rightarrow NO + I_2 + H_2O}$

Step 1: Assign oxidation number to all the atoms involved in the equation.

+1 +5 3(-2) +1 -1 +2 -2 0 2(+1)-2

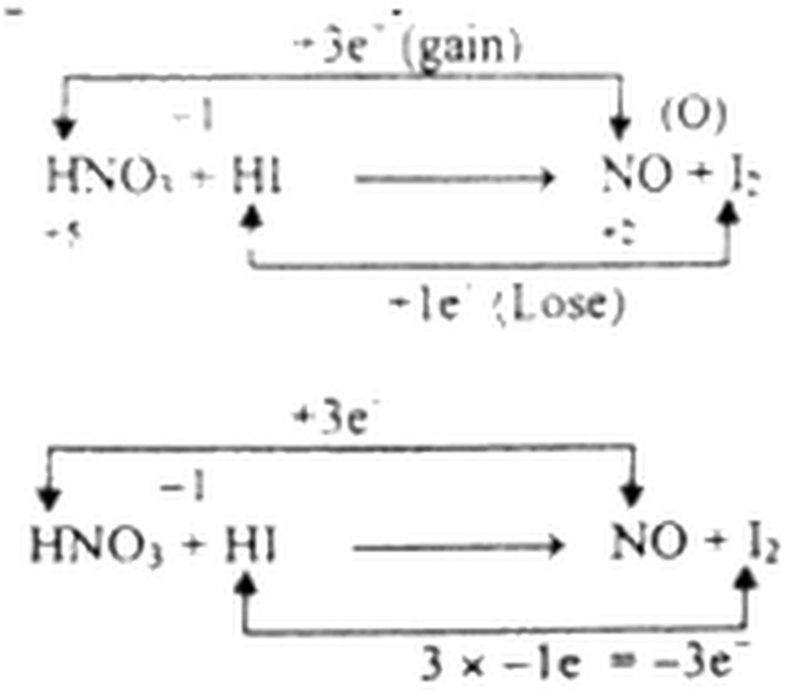
$$H\ N\ O_3 + H\ I \rightarrow N\ O + I_2 + H_2O$$

Step 2: Identify the elements undergoing a change in oxidation number

The I atom change their oxidation state from to -1 in HI to zero in I_2. This is due to loss of $1e^-$ per I atom. This is oxidation.

The N atom changes its oxidation state from +5 in HNO_3 to +2 in NO. This is due to gain of $3e^-$ per N atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. use multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$ gain by 2 and $2e^-$ loss by 3.

Multiply the coefficient of HI by 3 and that of I_2 by $\frac{3}{2}$, or better to multiply HI by 6 and I_2 by 3 in order to equalize the number of I atoms for electron loss.

$$HNO_3 + 6HI \rightarrow NO + 3I_2 + H_2O$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

To balance oxygen atoms multiply coefficient of HNO_3 by 2 and H_2O by 4.

$$2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$$

This is the balanced equation.

(iii) $Ag + H_2S + O_2 \rightarrow Ag_2S + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

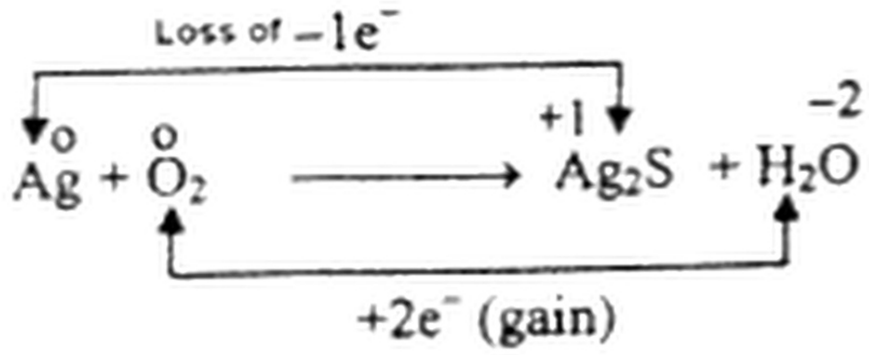
$$\overset{0}{Ag} + \overset{2(+1)\,-2}{H_2S} + \overset{0}{O_2} \rightarrow \overset{2(+1)\,-2}{Ag_2S} + \overset{2(+1)-2}{H_2O}$$

Step 2: Identify the elements undergoing a change in oxidation number

The Ag atom changes its oxidation state from zero in Ag to +1 in Ag_2S. This is due to loss of $1e^-$ per Ag atom. This is oxidation.

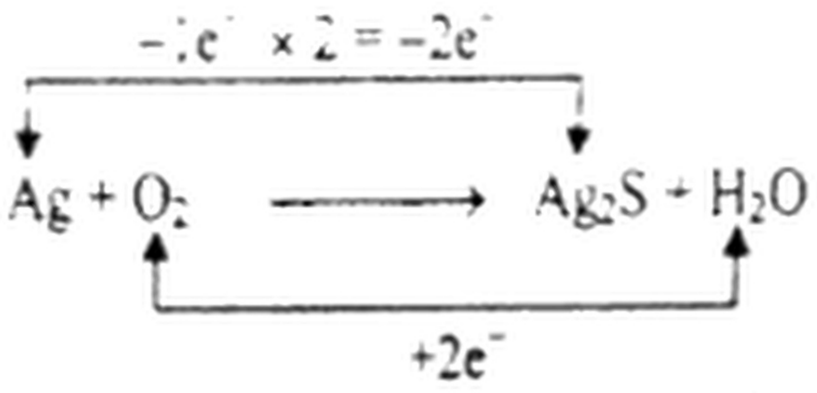
The O atom changes its oxidation state from 0 in O_2 to -2 in H_2O. This is due to gain of $2e^-$ per O atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $4e^-$ gain by 1 and ... loss by 4.



Multiply the coefficients of Ag by 4 and that of Ag_2S by 2,in order to equalize the number of Ag atoms for electrons loss.

$$4Ag + H_2S + O_2 \rightarrow 2Ag_2S + H_2O$$

Step 5: Balance the rest of the equation by inspection method.

To balance S atoms multiply coefficient of H_2S by 2.

$$4Ag + 2H_2S + O_2 \rightarrow 2Ag_2S + 2H_2O$$

This is the balanced equation.

(iv) $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

0		-1+5	3(-2)		+2	+5	2(-1)	+2 -2	2(+1) -2
Zn	+	HN	O_3	→	$Zn(NO_3)_2$	+	NO		$+H_2O$

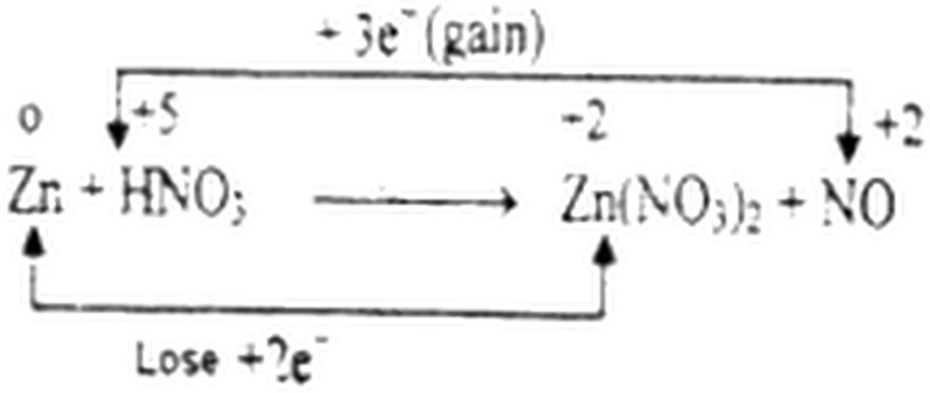
Step 2: Identify the elements undergoing a change in oxidation number

The Zn atoms change their oxidation state from zero in Zn to +2 in $Zn(NO_3)_2$. This is due to loss of $2e^-$ per Zn atom. This is oxidation.

Some N atoms change their oxidation state from+5 in $HNO_3$to +2 in NO. This is due to gain of $3e^-$ per N atom. This is reduction. Some N atoms retain their oxidation state at +5 in HNO_3 and $Zn(NO_3)_2$. Thus, write HNO_3 twice on L.H.S.

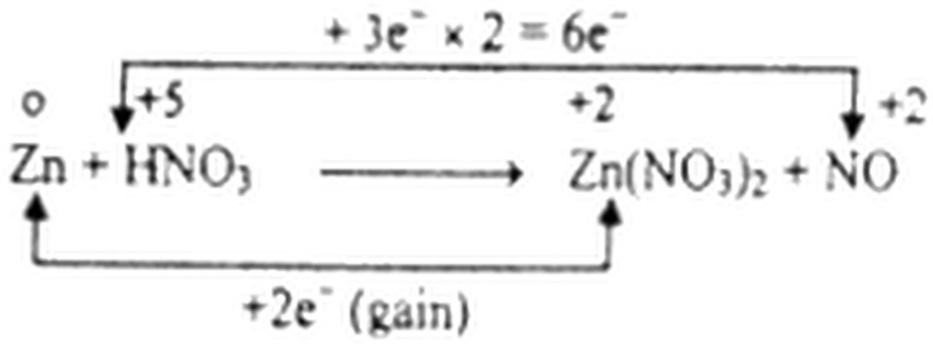
$$HNO_3 + Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO + H_2O$$

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$gain by 2 and $2e^-$loss by 3.



Multiply the coefficient of Zn and that of $Zn(NO_3)_2$ by 3.

Multiply the coefficient of NHO_3 and NO by 2.

$$HNO_3 + 3Zn + 2HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + H_2O$$

Step 5: Balance the rest of the equation by inspection method.

Balance equation in the following step-wise manner.

To balance nitrogen atoms multiply coefficient of other HNO_3 by 6.

To balance oxygen atoms multiply coefficient of H_2O by 4.

$$6HNO_3 + 3Zn + 2HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + 4H_2O$$

Or $$3Zn + 8HNO_3 \rightarrow 3Zn(NO_3)_2 + 2NO + 4H_2O$$

This is the balanced equation.

(v) $Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

0 2(+1)+6 4(-2) +2 +5 4(-2) +4 2(-2) 2(+1) -2

$Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

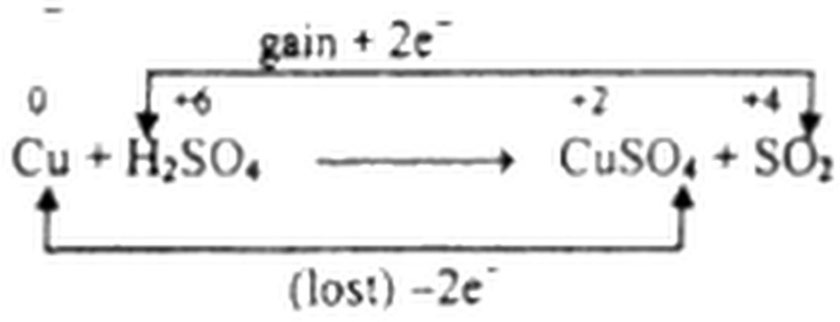
Step 2: Identify the elements undergoing a change in oxidation number

The Cu atoms change their oxidation state from zero in Cu to +2 in $CuSO_4$. This is due to loss of $2e^-$ per Cu atom. This is oxidation.

Some S atoms change their oxidation state from +6 in H_2SO_4 to +4 in SO_2. This is due to gain of $2e^-$ per S atom. This is reduction. Some S atoms retain their oxidation state at +6 in H_2SO_4 and $CuSO_4$. Thus, write H_2SO_4 twice on L.H.S.

$$H_2SO_4 + Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$$

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Since, loss and gain of electrons is equal, therefore balance the rest of the equation by inspection method.

To balance O atoms multiply coefficient of H_2O by 2.

$$H_2SO_4 + Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

Or

$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

This is the balanced equation.

(vi) $Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

0 +1 -2 +1 +1 -1 +1 +5 3(-2) 2(+1) -2

$Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$

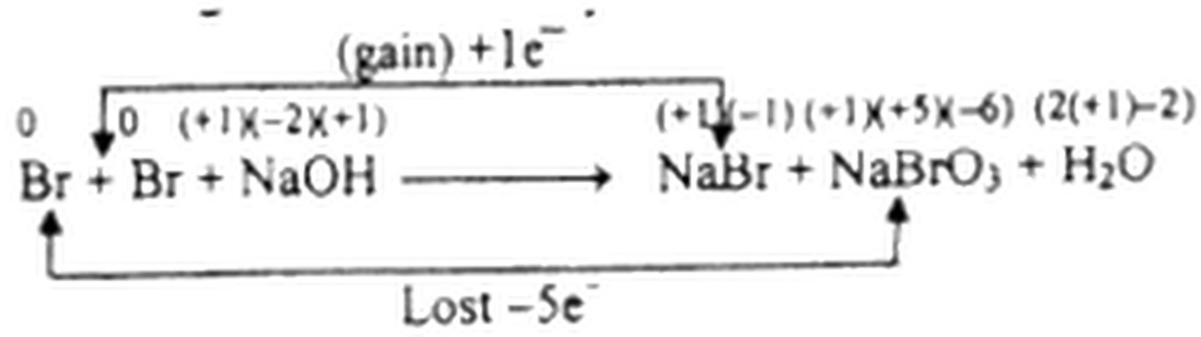
Step 2: Identify the elements undergoing a change in oxidation number

Some Br atoms change their oxidation state from zero in Br_2 to +5 in $NaBrO_3$. This is due to loss of $5e^-$ per Br atom. This is oxidation.

Some Br atoms change their oxidation state from zero in Br_2 to -1 in NaBr. This is due to gain of $1e^-$ per Br atom. This is reduction. So, write Br_2 twice on L.H.S.

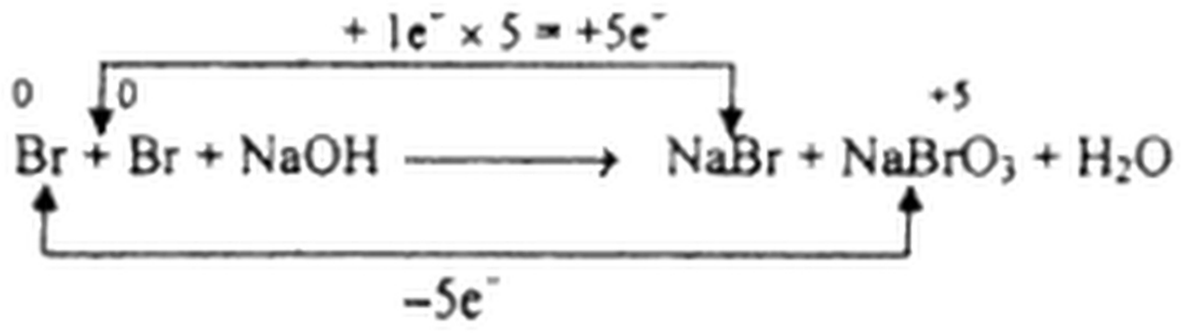
$$Br_2 + Br_2 + NaOH \rightarrow NaBr + NaBrO_3 + H_2O$$

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by 5 and $10e^-$ loss by 1.



Multiply the coefficient of Br_2 with 5 and that of NaBr with10 in order to equalize the number Br atoms for electron gain.

$$Br_2 + 5Br_2 + NaOH \rightarrow 10NaBr + NaBrO_3 + H_2O$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

Balance equation in the following step-wise manner.

To balance Br atoms multiply coefficient of $NaBrO_3$ by 2.

To balance sodium atoms multiply coefficient of NaOH by 12.

To balance oxygen atoms multiply coefficient of H_2O by 6

$$Br_2 + 5Br_2 + 12NaOH \rightarrow 10NaBr + 2NaBrO_3 + 6H_2O$$

Or $$6Br_2 + 12NaOH \rightarrow 10NaBr + 2NaBrO_3 + 6H_2O$$

Or divide by 2, $$3Br_2 + 6NaOH \rightarrow 5NaBr + NaBrO_3 + 3H_2O$$

This is the balanced equation.

(vii) $HI + H_2SO_4 \rightarrow I_2 + SO_4 + H_2O$

Step 1: Assign oxidation number to all the atoms involved in the equation.

+1 -1 2(+1)+6 4(-2) 0 +4 2(-2) 2(+1)-2

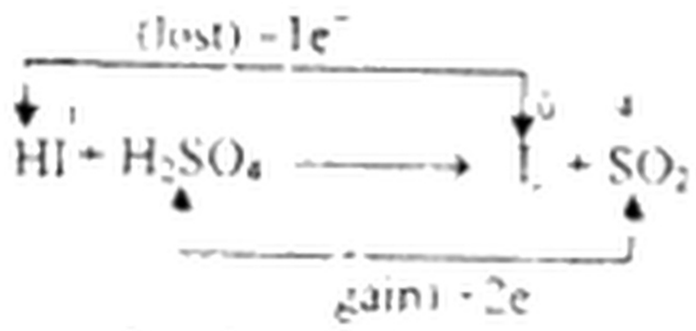
$HI + H_2SO_4 \rightarrow I_2 + SO_4 + H_2O$

Step 2: Identify the elements undergoing a change in oxidation number

The I atoms changes its oxidation state from -1 in HI to zero in I_2. This is due to loss of $1e^-$ per I atom. This is oxidation.

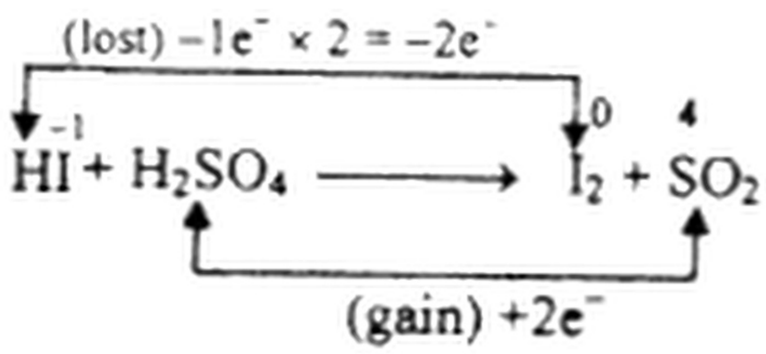
The S atom changes its oxidation state from +6 in H_2SO_4 to +4 in SO_2. This is due to gain of $2e^-$ per S atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $3e^-$ gain by 1 and $1e^-$ loss by 2.



Multiply the coefficients of HI by 2so that number of I atoms remains equal for electrons loss.

$$2HI + H_2SO_4 \rightarrow I_2 + SO_4 + H_2O$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

To balance oxygen atoms multiply coefficient of H_2O by 2.

$$2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$$

This is the balanced equation.

(viii) $NaClO_2 + Cl_2 \rightarrow NaCl + ClO_2$

Step 1: Assign oxidation number to all the atoms involved in the equation.

+1 +3 2(-2) 0 +1 -1 +4 2(-2)

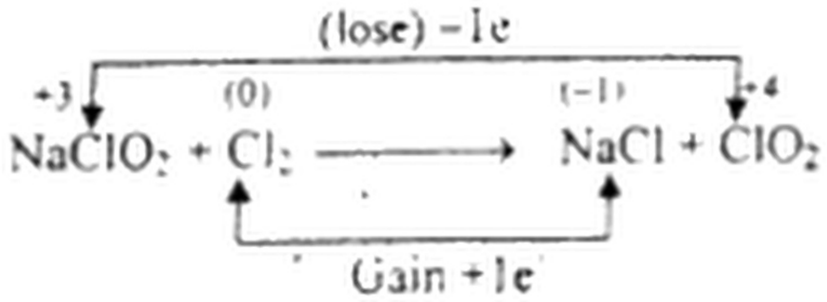
$NaClO_2 + Cl_2 \rightarrow NaCl + ClO_2$

Step 2: Identify the elements undergoing a change in oxidation number

The Cl atoms changes its oxidation state from +3 in $NaClO_2$ to +4 in ClO_2. This is due to loss of $1e^-$ per Cl atom. This is oxidation.

The Cl atom changes its oxidation state from zero in Cl_2 to -1 in NaCl. This is due to gain of $1e^-$ per Cl atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by 1 and $1e^-$ loss by 2.

Multiply the coefficients of $NaClO_2$ and that of ClO_2 by 2. so that number of Cl atoms remains equal for electrons loss.

$$2NaClO_2 + Cl_2 \rightarrow NaCl + 2ClO_2$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

To balance sodium atoms multiply coefficient of NaCl by 2.

$$2NaClO_2 + Cl_2 \rightarrow 2NaCl + 2ClO_2$$

This is the balanced equation.

(ix) $Na + H_2O \rightarrow NaOH + H_2$

Step 1: Assign oxidation number to all the atoms involved in the equation.

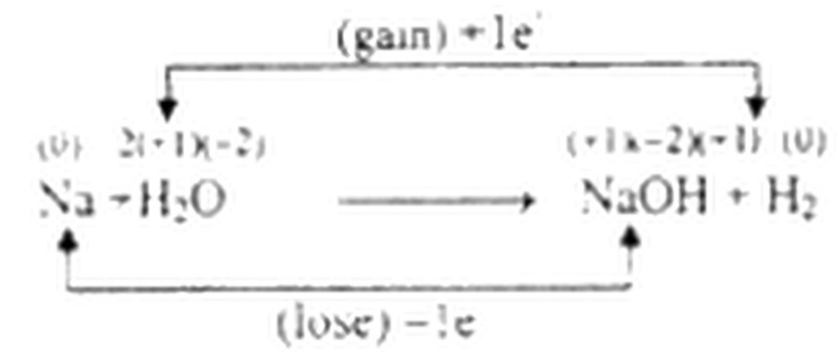
$$\begin{array}{ccccccccc} 0 & & 2(+1)-2 & & +1 & -2 & +1 & & 0 \\ Na & + & H_2O & \rightarrow & NaOH & & & + & H_2 \end{array}$$

Step 2: Identify the elements undergoing a change in oxidation number

The Na atom changes its oxidation state from zero in Na to +1 in $NaOH$. This is due to loss of $1e^-$ per Na atom. This is oxidation.

The H atom changes its oxidation state from +2 in H_2O to zero in H_2. This is due to gain of $1e^-$ per H atom. This is reduction.

Step 3: Note a change in oxidation state. Indicate this change by the number of electrons gained or lost by each element.



Step 4: Balance the loss and gain of electrons by multiplying with small suitable whole numbers. Use these multipliers as coefficient for respective substances.

Thus, to balance loss and gain of electrons in this reaction, multiply $2e^-$ gain by 1 and $1e^-$ loss by 2.

Multiply the coefficients of Na and that of $NaOH$ by 2. so that number of Na atoms equalize for electron loss.

Multiply coefficients of HNO_3 and NO by 2.

$$2Na + H_2O \rightarrow 2NaOH + H_2$$

These coefficients should not be changed after this.

Step 5: Balance the rest of the equation by inspection method.

To balance oxygen atoms multiply coefficient of H_2O by 2.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

This is the balanced equation.

16. **A dead Lead Storage battery has 5 g of $PbSO_4$ deposited on each of its anodes. It is connected to a charger that supplies 0.12 ampere of current at a voltage of13V. To which electrode, Pb or PbO_2 should the anode of charger be connected? Write the half reaction occurring at this electrode during charging?**

Ans: Recharging a dead lead storage battery:

i. Battery can be recharged by connecting its anode to the negative terminal of direct current and the cathode to the positive terminal of the direct current.

ii. Reverse chemical reactions occur at anode and cathode of the battery.

iii. Thus deposition of Pb on anode and PbO_2 on cathode takes place. In this way we can recharge the battery.

Redox reactions during recharging:

The reactions of recharging of battery are as follows.

At cathode:

$$PbSO_{4(s)} + 2e^- \rightarrow Pb_{(s)} + SO_4^{-2}{}_{(aq)}$$

At anode:

$$PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow PbO_{2(s)} + SO_4^{-2}{}_{(aq)} + 4H^+_{(aq)} + 2e^-$$

Net reaction: $$2PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq)} + 2SO_4^{-2}{}_{(aq)}$$

After recharging H_2SO_4 solution is concentrated again bringing density to its initial value of 1.25 g cm^{-3}

17. **Calculate the mass of Cu and O_2 produced by the electrolysis of $CuSO_4$ solution on passing 5.0 amperes of current for 2 hrs. What would be the volume of O_2 at S.T.P. following reactions occurs at the electrodes.**

At anode $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

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

(Ans: 0.1865g, 2.089 dm^3)

Solution:

Calculation of total charge passed:

Time = 2 hours = 2 × 60 × 60 = 7200 s

Current = 5 A

As we know that

$$current = \frac{charge}{Time(s)}$$

So charge = current × Time (s)

= 5 A × 7200 s

= 36000 C

As 96500 C = 1 F

$$36000\ C = \frac{1\ F}{96500\ C} \times 36000\ C = 0.3731\ F$$

In the electrolysis of molten bauxite, the cathode reaction is

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

It shows that for every 2 moles of electrons, 1 mole of Cu is produced.

1 F = 1 mole of electron.

2 F = 2 moles of electrons = 1 mole of Cu

Thus,

2 F charge produces Cu = 1 mole

0.3731 F charge produces Cu $= \frac{1}{2} \times 0.3731$

= 0.1866 moles

Atomic mass of Cu $= 63.5\ g\ mole^{-1}$

So 1 mole of Cu = 63.5 g

0.1866 moles of Cu = 63.5 × 0.1866

= 11.849 g

The reaction at anode is oxidation of O^2 ions.

$$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$$

It shows that for every 4 moles of electrons, 1 mole of O_2 is produced.

1F = 1 mole of electron

4F = 4 moles of electrons = 1 mole of O_2.

Thus,

4F electricity produce O_2 = 1 mole

0.3731 F of electricity produce O_2 $= \frac{1}{4} \times 0.3731$

$= 0.0933\ moles\ of\ O_2$

Molecular mass of O_2 $= 32\ g\ mole^{-1}$

Hence,

1 mole of O_2 = 32 g

0.0933 moles of Cu = 32 × 0.0933

= 2.986 g

Thus, at S.T.P

1 mole of O_2 = $22.414\ dm^3$

0.0933 moles of O_2 = $0.0933 \times 22.414\ dm^3$

= $2.091\ dm^3$

18. If a chromium steel bicycle handle bar is scratched, exposing steel, will chromium or steel corrode?

Ans: Chromium is ranked higher in activity series than iron (steel is a alloy of iron) so it has more activity than iron and it will react more radially with atmospheric oxygen hence the chromium will corrode. Thus if a chromium steel bicycle handle bar is scratched, exposing steel, will chromium will corrode instead of steel.

19. Use the activity series of metals to predict the products of following single replacement reactions, give reason.

i) $Cu(s) + HCI(aq) \rightarrow$

ii) $Fe_2O_3(s) + H_2(g) \rightarrow$

iii) $Mg(s) + CrCl_3(s) \rightarrow$

iv) $NiO(s) + H_2(g) \rightarrow$

v) $Cd(s) + CuSO_4(aq) \rightarrow$

Ans: In active series a metal which is ranked higher can displace Low ranked metal from its aqueous solution.

i) $Cu(s) + HCl(aq) \rightarrow$

Cu is ranked lower in activity series than hydrogen, thus has it will not displace hydrogen from HCl.

$Cu(s) + HCl(aq) \rightarrow No\ Reation$

ii) $Fe_2O_3(s) + H_2(g) \rightarrow$

The hydrogen is ranked below Fe in activity series, hence it cannot displace Fe.

$2Fe_2O_3(s) + 3H_2(g) \rightarrow No\ reaction$

iii) $Mg(s) + CrCl_3(s) \rightarrow$

Mg is ranked higher than Cr in activity series thus it can displace Cr thus this reaction will occur

$3Mg(s) + 2CrCl_3(s) \rightarrow 3MgCl_3 + 2Cr$

iv) $NiO(s) + H_2(g) \rightarrow$

The Hydrogen is ranked below Ni in activity series thus it cannot displace Ni

$NiO(s) + H_2(g) \rightarrow No\ reaction$

v) $Cd(s) + CuSO_4(aq) \rightarrow$

The Cd is ranked above Cu in activity series, hence it will display Cu. Thus this reaction will occur

$Cd(s) + CuSO_4(aq) \rightarrow CdSO_4 + Cu$

20. Justify that photosynthesis in plants is a redox reaction.

$$6CO_2 + 6H_2O \xrightarrow{chlorophyll} C_6H_{12}O_6 + 6CO_2$$

Solution:

Assigning oxidation numbers:

$$\overset{+4\ 2(-2)}{6CO_2} + \overset{2(+1)\ -2}{6H_2O} \xrightarrow{chlorophyll} \overset{6(0)\ 12(+1)\ 6(-2)}{C_6H_{12}O_6} + \overset{0}{6CO_2}$$

(i) Oxidation:

The O atom changes its Oxidation State from -2 to 0 in O_2.this is due to loss of electrons. Hence it is oxidation.

(ii) Reduction:

The C atom changes its Oxidation State from +4 to 0 in $C_6H_{12}O_6$ .this is due to gain of electrons. hence it is reduction.

Conclusion:

As oxidation and reduction occur in photosynthesis therefore it is a redox reaction

21. Metal A can displace metal B from its aqueous solution, metal B can displace metal D from its aqueous solution, but it cannot displace metal C from its aqueous solution. Metal C can displace metal A from its aqueous solution. Metal B and D can be displaced from their aqueous solution by hydrogen. Arrange these metals and hydrogen in order of decreasing activity.

Solution:

Given information:

i. Metal A can displace metal B from its aqueous solution. It means that metal A is more active than metal B.

$$\boldsymbol{Metal\ A} > Metal\ B$$

ii. Metal B can displace metal D from its aqueous solution. It means that metal B is more active than metal D.

$$\boldsymbol{Metal\ A} > Metal\ B > Metal\ D$$

iii. Metal C can displace metal A from its aqueous solution. It means that metal C is more active than metal A.

$$\boldsymbol{Metal\ C} > Metal\ A > Metal\ B > Metal\ D$$

iv. Metal B and D can be displaced by hydrogen from their aqueous solution. Thus, It means that metal B and D are less active than hydrogen.

Conclusion:

Arrange these metals and hydrogen in order of decreasing activity.

$$\boldsymbol{Metal\ C} > Metal\ A > Hydrogen > Metal\ B > Metal\ D$$

22. Differences in the standard potentials of electrodes A and B, B and C are 1.5 V and 3.0 V respectively.

a) Calculate the cell potential of the working cell between A and C.

b) Identify cathode in this cell.

Solution:

a) Calculate the cell potential of the working cell between A and C.

Given that

$E^\circ_B - E^\circ_A = 1.50V$ .....(i)

$E^\circ_B - E^\circ_C = 3.0V$ .....(ii)

For the calculate the cell potential of the working cell between A and C we have to subtract equation (i) from equation (ii)

$$(E^\circ_B - E^\circ_C) - (E^\circ_B - E^\circ_A) = 3.0 - 1.50$$

$$E^\circ_B - E^\circ_C - E^\circ_B + E^\circ_A = 1.50V$$

$$E^\circ_A - E^\circ_C = 1.50V$$

(b) Identify cathode in this cell.

$$E^\circ_A = 1.50V + E^\circ_C$$

As the standard electrode potential of A is greater than B therefore A will act as cathode and B will act as anode.

23. Standard electrode potentials of three electrodes are in the following order,

$$E^\circ_A > E^\circ_B < E^\circ_C$$

Identify anode in the working cell, when (a) Electrodes A and C are joined in series. (b) Electrodes B and C are joined in series (c) Electrodes A and B are joined in series.

Ans: The emf measured when a metal/metal ion electrode is coupled to a hydrogen electrode under standard conditions is know as the standard electrode potential of that metal/metal ion combination.

Condition:

As $E^\circ_A > E^\circ_B < E^\circ_C$

Principle:

The electrode which has higher standard electrode potential will act as cathode and the electrode which has lower standard electrode potential will act as anode.

The above equation indicates that standard electrode potential of electrode B is less than that of both A and C electrodes.

(a) Electrodes A and C are joined in series.

As no information is provided about the electrode potential of electrode A and C relative to one another. Therefore, we cannot identify anode in this case.

(b) Electrodes B and C are joined in series.

When electrodes B and C are joined in series electrode B will act as anode and electrode C will act as cathode. We conclude this with the help of following relation

Since, $E^\circ_B < E^\circ_C$

Therefore,

(c) Electrodes A and B are joined in series.

As we know that

$$E^\circ_A < E^\circ_B$$

Thus, when electrodes A and B are joined in series electrode B will act as anode and electrode will act as cathode because the electrode which has higher standard electrode potential will act as cathode and the electrode which has lower standard electrode potential will act as anode.

24. Evaluate the importance of standard reduction potentials.

Ans: i. A reduction potential measures the tendency of a molecule to be reduced by taking up new electrons.

ii. The standard reduction potential is the reduction potential of a molecule under specific, standard conditions.

iii. Standard reduction potentials can be useful in determining the directionality of a reaction.