NCERT SOLUTIONS

CLASS - 11th



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Class : 11th
Subject : Chemistry
Chapter : 8
Chapter Name : Redox Reactions

Q8.1 Assign oxidation number to the underlined elements in each of the following species:
$(a)\mathrm{NaH_2\underline{P}O_4}(b)\mathrm{NaH\underline{S}O_4}(c)\mathrm{H_4\underline{P}_2O_7}(d)\mathrm{K_2\underline{Mn}O_4}$
$(e)\mathrm{Ca\underline{O}_2}(f)\mathrm{Na\underline{B}H_4}(g)\mathrm{H_2\underline{S}_2O_7}$ (h) $\mathrm{KAl(\underline{S}O_4)_2 \cdot 12H_2O}$

Answer. (a) Let the oxidation number of P be x.
We know that,
Oxidation number of Na = +1
Oxidation number of H = +1
Oxidation number of O = –2
$\mathrm{NaH_2PO_4}$
Then, we have
$1(+1) + 2(+1) + 1(x) + 4(-2) = 0$
$\Rightarrow 1 + 2 + x - 8 = 0$
$\Rightarrow x = +5$
Hence, the oxidation number of P is +5.

(b) $\overset{+1}{Na}\ HSO_4$
Then, we have
$1(+1) + 1(+1) + 1(x) + 4(-2) = 0$
$\Rightarrow 1 + 1 + x - 8 = 0$
$\Rightarrow x = +6$
Hence, the oxidation number of S is + 6.

(c) $\mathrm{H_4P_2O_7}$
Then, we have
$4(+1) + 2(x) + 7(-2) = 0$
$\Rightarrow 4 + 2x - 14 = 0$
$\Rightarrow 2x = +10$
$\Rightarrow x = +5$
Hence, the oxidation number of P is + 5.

(d) $\mathrm{K_2MnO_4}$
Then, we have
$2(+1) + x + 4(-2) = 0$
$\Rightarrow 2 + x - 8 = 0$
$\Rightarrow x = +6$

Hence, the oxidation number of Mn is + 6.

(e) $\overset{+2 \quad x}{\mathrm{CaO_2}}$
Then, we have
$(+2) + 2(x) = 0$
$\Rightarrow 2 + 2x = 0$
$\Rightarrow x = -1$

(f) $9\mathrm{NaBH_4}$
Then, we have
$1(+1) + 1(x) + 4(-1) = 0$
$\Rightarrow 1 + x - 4 = 0$
$\Rightarrow x = +3$
Hence, the oxidation number of B is + 3.

(g) $\mathrm{H_2S_2O_7}$
Then, we have
$2(+1) + 2(x) + 7(-2) = 0$
$\Rightarrow 2 + 2x - 14 = 0$
$\Rightarrow 2x = 12$
$\Rightarrow x = +6$
Hence, the oxidation number of S is + 6.

(h) $\mathrm{KAl(\underline{SO}_4)_2 \cdot 12H_2O}$
Then, we have
$1(+1) + 1(+3) + 2(x) + 8(-2) + 24(+1) + 12(-2) = 0$
$\Rightarrow 1 + 3 + 2x - 16 + 24 - 24 = 0$
$\Rightarrow 2x = 12$
$\Rightarrow x = +6$
Or,
We can ignore the water molecule as it is a neutral molecule. Then, the sum of the oxidation numbers of all atoms of the water molecule may be taken as zero. Therefore, after ignoring the water molecule, we have
$1(+1) + 1(+3) + 2(x) + 8(-2) = 0$
$\Rightarrow 1 + 3 + 2x - 16 = 0$
$\Rightarrow 2x = 12$
$\Rightarrow x = +6$
Hence, the oxidation number of S is + 6.
Page : 280 , Block Name : Exercise

Q8.2 What are the oxidation numbers of the underlined elements in each of the following and how do you rationalise your results?
$\mathrm{Kl_3}$

$H_2\underline{S_4}O_6$

(c) Fe_3O_4

(d) CH_3CH_2OH

(e) CH_3COOH

Answer. (a) Kl_3

In Kl_3, the oxidation number (O.N.) of K is +1. Hence, the average oxidation number of I is $-\frac{1}{3}$. However, O.N. cannot be fractional. Therefore, we will have to consider the structure of KI3to find the oxidation states.

In a Kl_3 molecule, an atom of iodine forms a coordinate covalent bond with an iodine molecule.



Hence, in a Kl_3molecule, the O.N. of the two I atoms forming the l_2> molecule is 0, whereas the O.N. of the I atom forming the coordinate bond is as 1.

(b)

$H_2\underline{S_4}O_6$

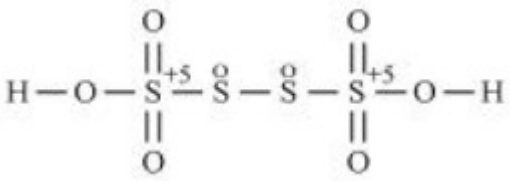
$\text{Now, } 2(+1) + 4(x) + 6(-2) = 0$

$\Rightarrow 2 + 4x - 12 = 0$

$\Rightarrow 4x = 10$

$\Rightarrow x = +2\frac{1}{2}$

However, O.N. cannot be fractional. Hence, S must be present in different oxidation states in the molecule.



The O.N. of two of the four S atoms is +5 and the O.N. of the other two S atoms is 0.

(c) Fe_3O_4

On taking the O.N. of O as 2, the O.N. of Fe is found to be $+2\frac{2}{3}$. However, O.N. cannot be fractional.

Here, one of the three Fe atoms exhibits the O.N. of +2 and the other two Fe atoms exhibit the O.N. of +3.

(d) CH_3CH_2OH

$2(x) + 6(+1) + 1(-2) = 0$

$\text{or, } 2x + 4 = 0$

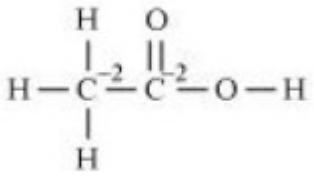
$\text{or, } x = -2$

$\text{Hence, the O.N. of } C \text{ is } \hat{a}\epsilon^u 2$

(e) CH_3COOH

$2(x) + 4(+1) + 2(-2) = 0$
$\text{or, } 2x = 0$
$\text{or, } x = 0$

However, 0 is average O.N. of C. The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number. Thus, C exhibits the oxidation states of +2 and â€“2 in CH_3COOH.



Page : 280 , Block Name : Exercise

Q8.3 Justify that the following reactions are redox reactions:
(a) $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$
(b) $Fe_2O_3(s) + 3cO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
(c) $4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3(s)$
(d) $2K(s) + F_2(g) \rightarrow 2K + F - (s)$
(e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

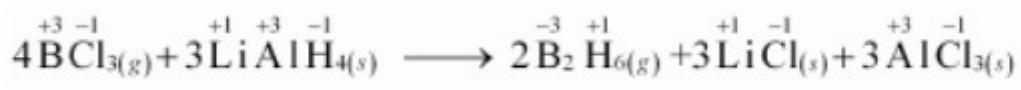
Answer. (a) $CuO_{(s)} + H_{2(g)}$longrightarrow $Cu_{(s)} + H_2O_{(g)}$
Let us write the oxidation number of each element involved in the given reaction as:
$\overset{+2}{Cu}\ O_{(s)} + \overset{\theta}{H}_{2(g)}$ -> $Cu_{(s)} + H_2O_{(g)}$
Here, the oxidation number of Cu decreases from +2 in CuO to 0 in Cu i.e., CuO is reduced to Cu. Also, the oxidation number of H increases from 0 in H_2 to +1 in H_2O i.e., H_2 is oxidized to H_2O. Hence, this reaction is a redox reaction.

(b) $Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$
Let us write the oxidation number of each element in the given reaction as:



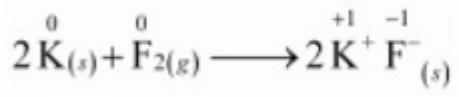
Here, the oxidation number of Fe decreases from +3 in Fe_2O_3 to 0 in Fe i.e., Fe_2O_3 is reduced to Fe. On the other hand, the oxidation number of C increases from +2 in CO to +4 in CO_2 i.e., CO is oxidized to CO_2. Hence, the given reaction is a redox reaction.

(c) $4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3(s)$
The oxidation number of each element in the given reaction can be represented as:



In this reaction, the oxidation number of B decreases from +3 in BCl_3 to −3 in B_2H_6. i.e., BCl_3 is reduced toB_2H_6. Also, the oxidation number of H increases from −1 in LiAlH4 to +1 in B_2H_6 i.e., LiAlH4 is oxidized to B_2H_6. Hence, the given reaction is a redox reaction.

(d) $2K(s) + F_2(g) \rightarrow 2K + F-$
The oxidation number of each element in the given reaction can be represented as:



(e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$
The oxidation number of each element in the given reaction can be represented as:

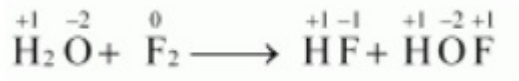


Here, the oxidation number of N increases from −3 in NH_3 to +2 in NO. On the other hand, the oxidation number of O_2 decreases from 0 in O_2 to −2 in NO and H_2O i.e., O_2 is reduced. Hence, the given reaction is a redox reaction.

Page : 280 , Block Name : Exercise

Q8.4 Fluorine reacts with ice and results in the change:
$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF(g)$
Justify that this reaction is a redox reaction:

Answer. Let us write the oxidation number of each atom involved in the given reaction above its symbol as:



Here, we have observed that the oxidation number of F increases from 0 in F_2 to +1 in HOF. Also, the oxidation number decreases from 0 in F_2 to −1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

Page : 280 , Block Name : Exercise

Q8.5 Calculate the oxidation number of sulphur, chromium and nitrogen in
H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^- Suggest structure of these compounds. Count for the fallacy.

Answer. (i) H_2SO_5

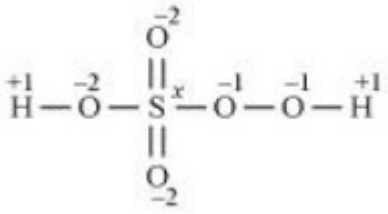
$2(+1) + 1(x) + 5(-2) = 0$

$\Rightarrow 2 + x - 10 = 0$

$\Rightarrow x = +8$

However, the O.N. of S cannot be +8. S has six valence electrons. Therefore, the O.N. of S cannot be more than +6.

The structure of H_2SO_5 is shown as follows:



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

$\Rightarrow 2 + x - 6 - 2 = 0$

$\Rightarrow x = +6$

Therefore, the O.N. of S is +6.

(ii)$Cr_2O_7^{2-}$

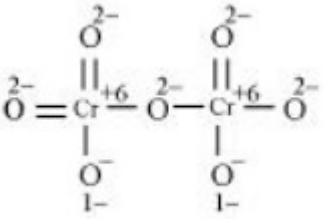
$2(x) + 7(-2) = -2$

$\Rightarrow 2x - 14 = -2$

$\Rightarrow x = +6$

Here, there is no fallacy about the O.N. of Cr in$Cr_2O_7^{2-}$

The structure of $Cr_2O_7^{2-}$ is shown as follows:



Here, each of the two Cr atoms exhibits the O.N. of +6.

(iii) NO_3^-

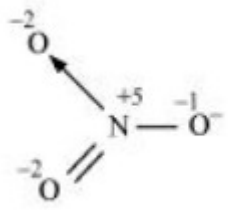
$1(x) + 3(-2) = -1$

$\Rightarrow x - 6 = -1$

$\Rightarrow x = +5$

Here, there is no fallacy about the O.N. of N in NO_3^-

The structure of NO_3^- is shown as follows:



The N atom exhibits the O.N. of +5.

Page : 280 , Block Name : Exercise

Q8.6 Write the formula for the following compounds:
(a) Mercury(II) chloride
(b) Nickel(II) sulphate
(c) Tin(IV) oxide
(d) Thallium(I) sulphate
(e) Iron(III) sulphate
(f) Chromium(III) oxide
Answer. (a) Mercury (II) chloride: $\mathrm{HgCl_2}$
(b) Nickel (II) sulphate: $\mathrm{NiSO_4}$
(c) Tin (IV) oxide: $\mathrm{SnO_2}$
(d) Thallium (I) sulphate: $\mathrm{TI_2SO_4}$
(e) Iron (III) sulphate: $\mathrm{Fe_2(SO_4)_3}$
(f) Chromium (III) oxide: $\mathrm{Cr_2O_3}$

Page : 280 , Block Name : Exercise

Q8.7 Suggest a list of the substances where carbon can exhibit oxidation states from −4 to +4 and nitrogen from −3 to +5.

Answer. The substances where carbon can exhibit oxidation states from −4 to +4 are listed in the following table.

Substance	O.N. of carbon
CH_2Cl_2	0
$ClC{\equiv}CCl$	+1
$HC{\equiv}CH$	–1
$CHCl_3$, CO	+2
CH_3Cl	–2
$Cl_3C - CCl_3$	+3
$H_3C - CH_3$	–3
CCl_4, CO_2	+4
CH_4	–4

The substances where nitrogen can exhibit oxidation states from –3 to +5 are listed in the following table.

Substance	O.N. of nitrogen
N_2	0
N_2O	+1
N_2H_2	−1
NO	+2
N_2H_4	−2
N_2O_3	+3
NH_3	−3
NO_2	+4
N_2O_5	+5

Page : 280 , Block Name : Exercise

Q8.8 While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

Answer. In sulphur dioxide (SO2), the oxidation number (O.N.) of S is +4 and the range of the O.N. that S can have is from +6 to −2.Therefore, SO2 can act as an oxidising as well as a reducing agent.
In hydrogen peroxide (H_2O_2) the O.N. of O is −1 and the range of the O.N. that O can have is from 0 to −2. O can sometimes also attain the oxidation numbers +1 and +2. Hence, H_2O_2 can act as an oxidising as well as a reducing agent.
In ozone , the O.N. of O is zero and the range of the O.N. that O can have is from 0 to −2. Therefore, the O.N. of O can only decrease in this case. Hence, O3 acts only as an oxidant.
In nitric acid (HNO_3), the O.N. of N is +5 and the range of the O.N. that N can have is from +5 to −3. Therefore, the O.N. of N can only decrease in this case. Hence, HNO_3 acts only as an oxidant.

Page : 280 , Block Name : Exercise

Q8.9 Consider the reactions:

(a) $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

(b) $O_3(g) + H_2O_2(l) \rightarrow H_2O(l) + 2O_2(g)$

Why it is more appropriate to write these reactions as:

(a) $6CO_2(g) + 12H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6H_2O(l) + 6O_2(g)$

(b) $O_3(g) + H_2O_2(I) \rightarrow H_2O(l) + O_2(g) + O_2(g)$

Also suggest a technique to investigate the path of the above (a) and (b) redox

Answer. (a)The process of photosynthesis involves two steps.

Step 1:

H_2O decomposes to give H_2 and O_2.

$2H_2O_{(l)} \longrightarrow 2H_{2(s)} + O_{2(s)}$

Step 2: TheH_2 produced in step 1 reduces CO_2, thereby producing glucose $(C_6H_{12}O_6)$ and H_2O.

$6CO_{2(g)} + 12H_{z(g)} \longrightarrow C_6H_{12}O_{6(s)} + 6H_2O_{(l)}$

Now, the net reaction of the process is given as:

$$\begin{aligned} &2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)} \Big] \times 6 \\ &6CO_{2(g)} + 12H_{2(g)} \longrightarrow C_6H_{12}O_{6(g)} + 6H_2O_{(l)} \\ \hline &6CO_{2(g)} + 12H_2O_{(l)} \longrightarrow C_6H_{12}O_{6(g)} + 6H_2O_{(l)} + 6O_{2(g)} \end{aligned}$$

It is more appropriate to write the reaction as given above because water molecules are also produced in the process of photosynthesis.

The path of this reaction can be investigated by using radioactive H_2O^{18} in place of H_2O.

(b) O_2 is produced from each of the two reactants O_3 and H_2O_2. For this reason, O_2 is written twice.

The given reaction involves two steps. First, O_3 decomposes to form O_2 and O. In the second step, H_2O_2 reacts with the O produced in the first step, thereby producing H_2O and O_2.

$$\begin{aligned} &O_{3(g)} \longrightarrow O_{2(g)} + O_{(g)} \\ &H_2O_{2(l)} + O_{(g)} \longrightarrow H_2O_{(l)} + O_{2(g)} \\ \hline &H_2O_{2(l)} + O_{3(g)} \longrightarrow H_2O_{(l)} + O_{2(g)} + O_{2(g)} \end{aligned}$$

The path of this reaction can be investigated by using $H_2O_2^{18}$or O_3^{18}.

Page : 280 , Block Name : Exercise

Q8.10 The compound AgF_2 is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

Answer. The oxidation state of Ag in AgF_2 is +2. But, +2 is an unstable oxidation state of Ag. Therefore, whenever AgF_2 is formed, silver readily accepts an electron to form Ag^+. This helps to bring the oxidation state of Ag down from +2 to a more stable state of +1. As a result, AgF_2 acts as a very strong oxidizing agent.

Page : 281 , Block Name : Exercise

Q8.11 Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

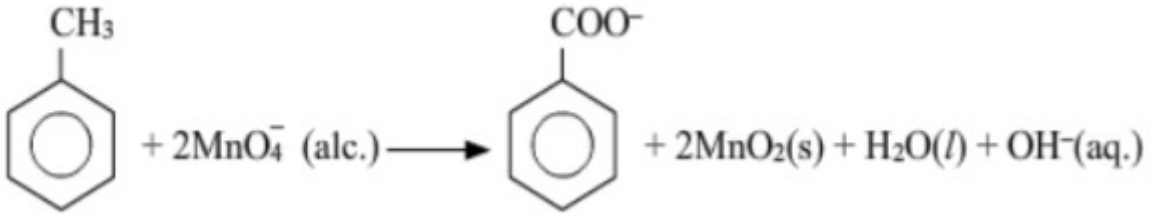
Answer. Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. This can be illustrated as follows:
(i) $\mathrm{P_4}$ and $\mathrm{F_2}$ are reducing and oxidising agents respectively.
If an excess of $\mathrm{P_4}$ is treated with $\mathrm{F_2}$, then $\mathrm{PF_3}$ will be produced, wherein the oxidation number (O.N.) of P is +3.
$\mathrm{P_4(\ excess\ ) + F_2 \longrightarrow PF_3}$
However, if P4 is treated with an excess of F2, then PF5 will be produced, wherein the O.N. of P is +5.
$\mathrm{P_4(\ excess\ ) + F_2 \longrightarrow PF_5}$
(ii)K acts as a reducing agent, whereas O2 is an oxidising agent.
If an excess of K reacts with O2, then K2O will be formed, wherein the O.N. of O is −2.
$\mathrm{4K(\ excess\ ) + O_2 \longrightarrow 2K_2O}$
However, if K reacts with an excess of O2, then K2O2 will be formed, wherein the O.N. of O is −1.
$\mathrm{2K + O_2(\ excess\ ) \longrightarrow K_2O_2}$
(iii)C is a reducing agent, while O2 acts as an oxidising agent.
If an excess of C is burnt in the presence of insufficient amount of O2, then CO will be produced, wherein the O.N. of C is +2.
$\mathrm{C(\ excess\ ) + O_2 \longrightarrow CO}$
On the other hand, if C is burnt in an excess of O2, then CO2 will be produced, wherein the O.N. of C is +4.
$\mathrm{C(\ excess\ ) + O_2 \longrightarrow CO}$

Page : 281 , Block Name : Exercise

Q8.12 How do you count for the following observations?
(a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.
(b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why?

Answer. (a) In the manufacture of benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant because of the following reasons.
(i) In a neutral medium, $\mathrm{OH^-}$ ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii) $KMnO_4$ and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol, $KMnO_4$ and toluene can react at a faster rate.
The balanced redox equation for the reaction in a neutral medium is give as below:



(b) When conc.H_2SO_4 is added to an inorganic mixture containing bromide, initially HBr is produced. HBr, being a strong reducing agent reduces H_2SO_4 to SO_2 with the evolution of red vapour of bromine.

$$2NaBr + 2H_2SO_4 \longrightarrow 2NaHSO_4 + 2HBr$$

$$2HBr + H_2SO_4 \longrightarrow Br_2 + SO_2 + 2H_2O$$

But, when conc. H_2SO_4 is added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, being a weak reducing agent, cannot reduce H_2SO_4 to SO_2.

$$2NaCl + 2H_2SO_4 \longrightarrow 2NaHSO_4 + 2HCl\,.$$

Page : 281 , Block Name : Exercise

Q8.13 Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:

(a) $2AgBr(s) + C_6H_6O_2(aq) \rightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2(aq)$

(b) $HCHO(l) + 2\left[Ag(NH_3)_2\right]^+(aq) + 3OH^-(aq) \rightarrow 2Ag(s) + HCOO^-(aq) + 4NH_3(aq) + 2H_2O(l)$

(c) $HCHO\ (1) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow Cu_2O(s) + HCOO^-(aq) + 3H_2O(l)$

(d) $N_2H_4(l) + 2H_2O_2(l) \rightarrow N_2(g) + 4H_2O(l)$

(e) $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

Answer. (a) Oxidised substance → $C_6H_6O_2$
Reduced substance → AgBr
Oxidising agent → AgBr
Reducing agent → $C_6H_6O_2$
(b)Oxidised substance → HCHO
Reduced substance → $\left[Ag\,(NH_3)_2\right]^+$
Oxidising agent → $\left[Ag\,(NH_3)_2\right]^+$
Reducing agent → HCHO
(c) Oxidised substance → HCHO
Reduced substance → Cu^{2+}

Oxidising agent → Cu^{2+}
Reducing agent → HCHO
(d) Oxidised substance → N_2H_4
Reduced substance → H_2O_2
Oxidising agent → H_2O_2
Reducing agent → N_2H_4
(e) Oxidised substance → Pb
Reduced substance → PbO_2
Oxidising agent → PbO_2
Reducing agent → Pb

Page : 281 , Block Name : Exercise

Q8.14 Consider the reactions:
$2S_2O_3^{2-}(aq) + I_2(s) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$
$S_2O_3^{2-}aq) + 2Br_2(l) + 5H_2O(l) \rightarrow^{2SO_4^{2-}} (aq) + 4Br^-(aq) + 10H^+$

Answer. The average oxidation number (O.N.) of S in $S_2O_3^{2-}$ is +2. Being a stronger oxidising agent than I2, Br2 oxidises $S_2O_3^{2-}$ to SO_4^{2-} , in which the O.N. of S is +6. However, I2 is a weak oxidising agent. Therefore, it oxidises $S_2O_3^{2-}$ to
$S_4O_6^{2-}$, in which the average O.N. of S is only +2.5. As a result, $S_2O_3^{2-}$ reacts differently with iodine and bromine.

Page : 281 , Block Name : Exercise

Q8.15 Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.

Answer. F_2 can oxidize Cl^- to Cl_2, Br^- to Br_2, and I^- to I_2
On the other hand, $Cl_2, Br_2,$ and I_2 cannot oxidize F^- to F_2. The oxidizing power of halogens increases in the order of$I_2 < Br_2 < Cl_2 < F_2$. Hence, fluorine is the best oxidant among halogens.
HI and HBr can reduce H_2SO_4 to SO_2, but HCl and HF cannot. Therefore, HI and HBr are stronger reductants than HCl and HF.
$2HI + H_2SO_4 \longrightarrow I_2 + SO_2 + 2H_2O$
$2HBr + H_2SO_4 \longrightarrow Br_2 + SO_2 + 2H_2O$
1^- can reduce Cu^{2+} to $Cu^+,$ but Br^-
$4I^-_{(aq)} + 2Cu^{2+}_{(aq)} \longrightarrow Cu_2I_{2(s)} + I_{2(aq)}$
Hence, hydroiodic acid is the best reductant among hydrohalic compounds.
Thus, the reducing power of hydrohalic acids increases in the order of HF < HCl < HBr < HI.

Page : 281 , Block Name : Exercise

Q8.16 Why does the following reaction occur?
$XeO_6^{4-}(aq) + 2F^-(aq) + 6H^+(aq) \rightarrow XeO_3(g) + F_2(g) + 3H_2O(l)$

What conclusion about the compound Na_4XeO_6(of which XeO_6^{4-} is a part) can be drawn from the reaction.

Answer. The given reaction occurs because XeO_6^{4-} oxidises $\mathbf{F}^-$ and $\mathbf{F}^-$reduces XeO_6^{4-}.

$XeO_{6(aq)}^{+} + 2F_{(aq)}^{-1} + 6H_{(aq)}^{+}$ -> $X_{3(g)} + F_{2(x)} + \overset{0}{3}H_2O_{(l)}$

In this reaction, the oxidation number (O.N.) of Xe decreases from +8 in XeO_6^{4-} to +6 in XeO3 and the O.N. of F increases from –1 in F^- to O in F_2 Hence, we can conclude that $Na_4XeO_6$is a stronger oxidising agent than F^-.

Page : 281 , Block Name : Exercise

Q8.17 consider the reactions:

$(a)\ H_3PO_2(aq) + 4AgNO_3(aq) + 2H_2O(l) \rightarrow H_3PO_4(aq) + 4Ag(s) + 4HNO_3(aq)$

$(b)\ H_3PO_2(aq) + 2CuSO_4(aq) + 2H_2O(I) \rightarrow H_3PO_4(aq) + 2Cu(s) + H_2SO_4(aq)$

$(c)\ C_6H_5CHO(I) + 2\left[Ag(NH_3)_2\right]^+(aq) + 3OH^-(aq) \rightarrow C_6H_5COO^-(aq) + 2Ag(s) + 4NH_3(aq) + 2H_2O(l)$

$(d)\ C_6H_5CHO(l) + 2Cu^{2+(aq)} + 5OH^-(aq) \rightarrow$ No change observed.

Answer. Cu^{2+} and Cu^{2+} act as oxidising agents in reactions (a) and (b)respectively.
In reaction (c), Cu^{2+} oxidises C_6H_5CHO to $C_6H_5COO^-$, but in reaction (d), Cu^{2+} cannot oxidise C_6H_5CHO.
Hence, we can say that Cu^{2+} is a stronger oxidising agent than Cu^{2+}.

Page : 281 , Block Name : Exercise

Q8.18 Balance the following redox reactions by ion-electron method:

$MnO_4^-(aq) + I^-(aq) \rightarrow MnO_2(s) + I_2(s)(\text{ in basic medium})$

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

$(c)\ H_2O_2(aq) + Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + H_2O(l)$ (in acidic solution)

$Cr_2O_7^{2-} + SO_2(g) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)(\text{ in acidic solution })$

Answer. (a) Step 1: The two half reactions involved in the given reaction are:

Oxidation half reaction: $\overset{-1}{I}_{(aq)} \longrightarrow I_{2(s)}$

Reduction half reaction: $\overset{+7}{M}\ nO_{4(aq)}^- \longrightarrow \overset{+4}{M}\ nO_{2(aq)}$

Step 2: Balancing I in the oxidation half reaction, we have:

$2I_{(aq)}^- \longrightarrow I_{2(s)}$

Now, to balance the charge, we add 2 e– to the RHS of the reaction.
$2I^-_{(aq)} \longrightarrow I_{2(x)} + 2e^-$
Step 3: In the reduction half reaction, the oxidation state of Mn has reduced from +7 to +4. Thus, 3 electrons are added to the LHS of the reaction.
$MnO^-_{4(aq)} + 3e^- \longrightarrow MnO_{2(aq)}$
Now, to balance the charge, we add 4 OH– ions to the RHS of the reaction as the reaction is taking place in a basic medium.
$MnO^-_{4(oq)} + 3e^- \longrightarrow MnO_{2(cq)} + 4OH^-$
Step 4: In this equation, there are 6 O atoms on the RHS and 4 O atoms on the LHS. Therefore, two water molecules are added to the LHS.
$MnO^-_{4(cq)} + 2H_2O + 3e^- \longrightarrow MnO_{2(ca)} + 4OH^-$
Step 5: Equalising the number of electrons by multiplying the oxidation half reaction by 3 and the reduction half reaction by 2, we have:
$6I^-_{(aq)} \longrightarrow 3I_{2(s)} + 6e^-$
$2MnO^-_{4(aq)} + 4H_2O + 6e^- \longrightarrow 2MnO_{2(s)} + 8OH^-_{(aq)}$
Step 6: Adding the two half reactions, we have the net balanced redox reaction as:
$6I^-_{(eq)} + 2MnO^-_{4(cy)} + 4H_2O_{(l)}$ -> $3I_{2(s)} + 2MnO_{2(s)} + 8OH^-_{(aq)}$

(b)Following the steps as in part (a), we have the oxidation half reaction as:
$SO_{2(x)} + 2H_2O_{(l)}$
And the reduction half reaction as:
$MnO^-_{4(\alpha q)} + 8H^+_{(aq)} + 5e^-$ -> $Mn^{2+}_{(aq)} + 4H_2O_{(l)}$
Multiplying the oxidation half reaction by 5 and the reduction half reaction by 2, and then by adding them, we have the net balanced redox reaction as:
$2MnO^-_{4(aq)} + 5SO_{2(g)} + 2H_2O_{(l)} + H^+_{(aq)}$ -> $2Mn^{2+}_{(aq)} + 5HSO^-_{4(aq)}$

(c) Following the steps as in part (a), we have the oxidation half reaction as:
$Fe^{2+}_{(\alpha q)} \longrightarrow Fe^{3+}_{(aq)} + e^-$
And the reduction half reaction as:
$H_2O_{2(aq)} + 2H^+_{(aq)} + 2e^- \longrightarrow 2H_2O_{(l)}$
Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:
$H_2O_{2(aq)} + 2Fe^{2+}_{(aq)} + 2H^+_{(aq)}$ -> $2Fe^{3+}_{(aq)} + 2H_2O_{(l)}$

(d) Following the steps as in part (a), we have the oxidation half reaction as:
$SO_{2(g)} + 2H_2O_{l)}$ -> $SO^{2-}_4 + 4H^+_{(aq)} + 2e^-$
And the reduction half reaction as:
$Cr_2O^{2-}_7 + 14H^+_{(aq)} + 6e^-$ -> $2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$
Multiplying the oxidation half reaction by 3 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:
$Cr_2O^{2-}_7 + 3SO_{2(g)} + 2H^+_{(aq)}$ -> $2Cr^{3+}_{(aq)} + 3SO^{2-}_4 + H_2O_{(l)}$

Page : 282 , Block Name : Exercise

Q8.19 Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

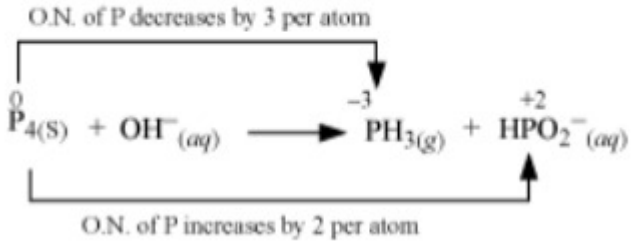
$P_{4(s)} + OH_{-(oq)}$ -> $PH_{3(g)} + HPO_2^-$

$N_2H_{4(l)} + ClO_{aq}^-$ -> $NO_{(g)} + Cl_{(g)}^-$

$Cl_2O_{7(g)} + H_2O_{2(aq)}$ -> $ClO_{2(aq)}^- + O_{2(g)}$ +$H_{(aq)}^+$

Answer. (a)

(a)



The O.N. (oxidation number) of P decreases from 0 in P_4 to −3 in PH3 and increases from 0 inP_4 to + 2 inHPO_2^- . Hence, P4acts both as an oxidizing agent and a reducing agent in this reaction.

Ion−electron method:

The oxidation half equation is:

$P4s \rightarrow HPO2 - (aq)$

The P atom is balanced as:

P_4 →4HPO2- (aq)

The O atom is balanced by adding 8 H2O molecules:

P_4 +8H_2O →4(aq)

The H atom is balanced by adding 12 H+ ions:

P_4 +8H_2O→4$HPO_2^-(aq) + 12H+$

The charge is balanced by adding e− as:

P_4 +8H_2O →4$HPO_2^-(aq) + 12H + +8e-$ ...(i)

The reduction half equation is:

$P_{4(s)} \longrightarrow PH_{3(g)}$

The P atom is balanced as:

P_4→4$PH_{3(g)}$

The H is balanced by adding 12 H+ as:

P_4 +12H+→4$PH_{3(g)}$

The charge is balanced by adding 12e− as:

P_4 +12H++12e-→4$PH_{3(g)}$ ...(ii)

By multiplying equation (i) with 3 and (ii) with 2 and then adding them, the balanced chemical equation can be obtained as:

5P_4 (s) +24H_2O→12HPO_2^-

− +8$PH_{3(g)} + 12H+$

As, the medium is basic, add 12OH− both sides as:

5P_4 (s) +12H_2O +12OH- →12HPO_2^- +8$PH_{3(g)}$

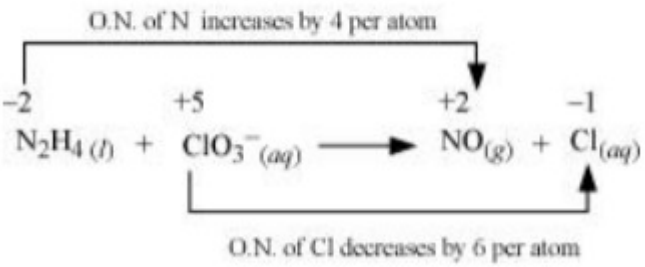
This is the required balanced equation.

Oxidation number method:
Let, total no of P reduced = x
∴Total no of P oxidised = 4– x
P_4 (s) +OH- $\rightarrow xPH_{3(g)} + 4\text{-}xHHPO_2^-$ ... (i)
Total decrease in oxidation number of P = x × 3 = 3x
Total increase in oxidation number of P = (4 – x) × 2 = 8 – 2x
∵ 3x = 8 – 2x x = 8/5 From (i),
5P4 (s) +5OH- $\rightarrow 8PH_{3(g)}) + 12HPO_2^-$

Since, reaction occurs in basic medium, the charge is balanced by adding 7OH– on LHS as:
$5P_4$ (s) +12OH- $\rightarrow 8PH_{3(g)} + 12HPO_2^-$
The O atoms are balanced by adding $12H_2O$ as:
$5P_4$ (s) + $12H_2O$ + 12OH- $\rightarrow$ $+12HPO_2^-$ + $8PH_{3(g)}$
This is the required balanced equation.

(b)

(b)



The oxidation number of N increases from – 2 in N_2H_4 to + 2 in NO and the oxidation number of Cl decreases from + 5 in ClO_3^- to – 1 in Cl^-. Hence, in this reaction, N_2H_4 is the reducing agent and ClO_3^- is the oxidizing agent.
Ion–electron method:

The oxidation half equation is:

$$\overset{-2}{N_2}H_{4(l)} \longrightarrow \overset{+2}{N}O_{(g)}$$

The N atoms are balanced as:

$$N_2H_{4(l)} \longrightarrow 2NO_{(g)}$$

The oxidation number is balanced by adding 8 electrons as:

$$N_2H_{4(l)} \longrightarrow 2NO_{(g)} + 8e^-$$

The charge is balanced by adding 8 OH^-ions as:

$$N_2H_{4(l)} + 8OH^-_{(aq)} \longrightarrow 2NO_{(g)} + 8e^-$$

The O atoms are balanced by adding $6H_2O$ as:

$$N_2H_{4(l)} + 8OH^-_{(aq)} \longrightarrow 2NO_{(g)} + 6H_2O_{(l)} + 8e^- \quad .....(i)$$

The reduction half equation is:

$$\overset{+5}{Cl}O^-_{3(aq)} \longrightarrow \overset{-1}{Cl}^-_{(aq)}$$

The oxidation number is balanced by adding 6 electrons as:

$$ClO^-_{3(aq)} + 6e^- \longrightarrow Cl^-_{(aq)}$$

The charge is balanced by adding $6OH^-$ ions as:

$$ClO^-_{3(aq)} + 6e^- \longrightarrow Cl^-_{(aq)} + 6OH^-_{(aq)}$$

The O atoms are balanced by adding $3H_2O$ as:

$$ClO^-_{3(aq)} + 3H_2O_{(l)} + 6e^- \longrightarrow Cl^-_{(aq)} + 6OH^-_{(aq)} \quad .....(ii)$$

The balanced equation can be obtained by multiplying equation (i) with 3 and equation (ii) with 4 and then adding them as:

$$3N_2H_{4(l)} + 4ClO^-_{3(aq)} \longrightarrow 6NO_{(g)} + 4Cl^-_{(aq)} + 6H_2O_{(l)}$$

Oxidation number method:

Total decrease in oxidation number of N = 2 × 4 = 8

Total increase in oxidation number of Cl = 1 × 6 = 6

On multiplying N_2H_4 with 3 and ClO^-_3 with 4 to balance the increase and decrease in O.N., we get:

$$3N_2H_{4(l)} + 4ClO^-_{3(aq)} \longrightarrow NO_{(g)} + Cl^-_{(aq)}$$

The N and Cl atoms are balanced as:

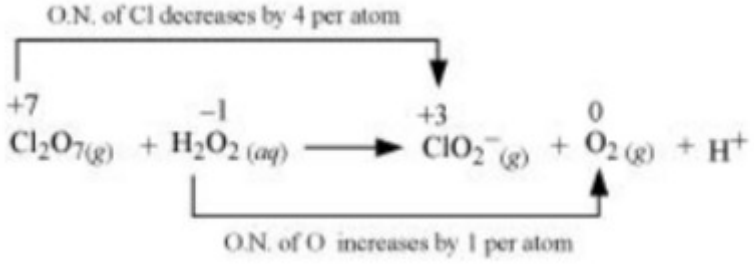
$$3N_2H_{4(l)} + 4ClO^-_{3(aq)} \longrightarrow 6NO_{(g)} + 4Cl^-_{(aq)}$$

The O atoms are balanced by adding $6H_2O$ as:

$$3N_2H_{4(l)} + 4ClO^-_{3(aq)} \longrightarrow 6NO_{(g)} + 4Cl^-_{(aq)} + 6H_2O_{(l)}$$

This is the required balanced equation.

(c)



The oxidation number of Cl decreases from + 7 in Cl_2O_7 to + 3 in ClO_2^- and the oxidation number of O increases from – 1 in H_2O_2 to zero in O_2. Hence, in this reaction, Cl_2O_7 is the oxidizing agent and H_2O_2 is the reducing agent.
Ion–electron method:

The oxidation half equation is:

$$\overset{-1}{H_2O_{2(aq)}} \longrightarrow \overset{0}{O_{2(g)}}$$

The oxidation number is balanced by adding 2 electrons as:

$$H_2O_{2(aq)} \longrightarrow O_{2(g)} + 2e^-$$

The charge is balanced by adding $2OH^-$ ions as:

$$H_2O_{2(aq)} + 2OH^-_{(aq)} \longrightarrow O_{2(g)} + 2e^-$$

The oxygen atoms are balanced by adding $2H_2O$ as:

$$H_2O_{2(aq)} + 2OH^-_{(aq)} \longrightarrow O_{2(g)} + 2H_2O_{(l)} + 2e^- \qquad \text{(i)}$$

The reduction half equation is:

$$\overset{+7}{Cl_2O_{7(g)}} \longrightarrow \overset{+3}{ClO^-_{2(aq)}}$$

The Cl atoms are balanced as:

$$Cl_2O_{7(g)} \longrightarrow 2ClO^-_{2(aq)}$$

The oxidation number is balanced by adding 8 electrons as:

The oxidation number is balanced by adding 8 electrons as:

$$Cl_2O_{7(g)} + 8e^- \longrightarrow 2ClO^-_{2(aq)}$$

The charge is balanced by adding $6OH^-$ as:

$$Cl_2O_{7(g)} + 8e^- \longrightarrow 2ClO^-_{2(aq)} + 6OH^-_{(aq)}$$

The oxygen atoms are balanced by adding $3H_2O$ as:

$$Cl_2O_{7(g)} + 3H_2O_{(l)} 8e^- \longrightarrow 2ClO^-_{2(aq)} + 6OH^-_{(aq)} \qquad \text{(ii)}$$

The balanced equation can be obtained by multiplying equation (i) with 4 and adding equation (ii) to it as:

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} + 2OH^-_{(aq)} \longrightarrow 2ClO^-_{2(aq)} + 4O_{2(g)} + 5H_2O_{(l)}$$

Oxidation number method:

Total decrease in oxidation number of $Cl_2O_7 = 4 \times 2 = 8$

Total increase in oxidation number of $H_2O_2 = 2 \times 1 = 2$

By multiplying H_2O_2 and O_2 with 4 to balance the increase and decrease in the oxidation number, we get:

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \longrightarrow ClO^-_{2(aq)} + 4O_{2(g)}$$

The Cl atoms are balanced as:

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \longrightarrow 2ClO^-_{2(aq)} + 4O_{2(g)}$$

The O atoms are balanced by adding $3H_2O$ as:

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \longrightarrow 2ClO^-_{2(aq)} + 4O_{2(g)} + 3H_2O_{(l)}$$

The H atoms are balanced by adding $2OH^-$ and $2H_2O$ as:

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} + 2OH^-_{(aq)} \longrightarrow 2ClO^-_{2(aq)} + 4O_{2(g)} + 5H_2O_{(l)}$$

This is the required balanced equation.

Page : 282 , Block Name : Exercise

Q8.20 What sorts of informations can you draw from the following reaction ?
$(CN)_{2(g)} + 2OH^-_{(aq)}$ -> $CN^-_{(g)} + CNO^-_{(aq)} + H_2O_{(l)}$

Answer. The oxidation numbers of carbon in $(CN)^2$, CN^- and CNO^- are +3, +2 and +4 respectively. These are obtained as shown below:
Let the oxidation number of C be x.
$(CN)_2$
2(x – 3) = 0
∴x = 3
CN^-
x – 3 = –1
∴x = 2
CNO–
x – 3 – 2 = –1
∴x = 4
The oxidation number of carbon in the various species is:

$$\left(\overset{+3}{C}N\right)_{2(g)} + 2OH^-_{(aq)} \longrightarrow \overset{+2}{C}N^-_{(aq)} + \overset{+4}{C}NO^-_{(aq)} + H_2O_{(l)}$$

It can be easily observed that the same compound is being reduced and oxidised simultaneously in the given equation. Reactions in which the same compound is reduced and oxidised is known as disproportionation reactions. Thus, it can be said that the alkaline decomposition of cyanogen is an example of disproportionation reaction.

Page : 282 , Block Name : Exercise

Q8.21 The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+}, MnO_2 and H^+ ion. Write a balanced ionic equation for the reaction.

Answer. The given reaction can be represented as:
$Mn^{3+}_{(aq)} \longrightarrow Mn^{2+}_{(aq)} + MnO_{2(s)} + H^+_{(aq)}$
The oxidation half equation is:
$Mn^{3+}_{(aq)} \longrightarrow \overset{+4}{M}nO_{2(s)}$
The oxidation number is balanced by adding one electron as:
$Mn^{3+}_{(aq)} \longrightarrow MnO_{2(s)} + e^-$
The charge is balanced by adding 4H+ ions as:
$Mn^{3+}_{(\alpha)} \longrightarrow MnO_{2(s)} + 4H^+_{(aq)} + e^-$
The O atoms and H+ ions are balanced by adding $2H_2O$ molecules as:

$$Mn^{3+}_{(aq)} + 2H_2O_{(l)} \longrightarrow MnO_{2(s)} + 4H^+_{(aq)} + e^- \ldots\ldots(i)$$

The reduction half equation is:

$Mn^{3+}_{(aq)} \longrightarrow Mn^{2+}_{(aq)}$
The oxidation number is balanced by adding one electron as:

$$Mn^{3+}_{(aq)} + e^- \longrightarrow Mn^{2+}_{(aq)} \quad .....(ii)$$

The balanced chemical equation can be obtained by adding equation (i) and (ii) as:

$$2Mn^{3+}_{(aq)} + 2H_2O_{(l)} \longrightarrow MnO_{2(s)} + 2Mn^{2+}_{(aq)} + 4H^+_{(aq)}$$

Page : 282 , Block Name : Exercise

Q8.22 Consider the elements:
Cs, Ne, I and F
(a) Identify the element that exhibits only negative oxidation state.
(b) Identify the element that exhibits only positive oxidation state.
(c) Identify the element that exhibits both positive and negative oxidation states.
(d) Identify the element which exhibits neither the negative nor does the positive oxidation state.

Answer. (a) F exhibits only negative oxidation state of −1.
(b) Cs exhibits positive oxidation state of +1.
(c) I exhibits both positive and negative oxidation states. It exhibits oxidation states of − 1, + 1, + 3, + 5, and + 7.
(d) The oxidation state of Ne is zero. It exhibits neither negative nor positive oxidation states.

Page : 282 , Block Name : Exercise

Q8.23 Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Answer. The given redox reaction can be represented as:
$Cl_{2(s)} + SO_{2(aq)} + H_2O_{(l)}$ -> $Cl^-_{(aq)} + SO^{2-}_4$
The oxidation half reaction is.

$$\overset{+4}{S}O_{2(aq)} \longrightarrow \overset{+6}{S}O_4^{2-}{}_{(aq)}$$

The oxidation number is balanced by adding two electrons as:

$$SO_{2(aq)} \longrightarrow SO_4^{2-}{}_{(aq)} + 2e^-$$

The charge is balanced by adding $4H^+$ ions as:

$$SO_{2(aq)} \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2e^-$$

The O atoms and H^+ ions are balanced by adding $2H_2O$ molecules as:

$$SO_{2(aq)} + 2H_2O_{(l)} \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2e^- \quad .....(i)$$

The reduction half reaction is:

$$Cl_{2(s)} \longrightarrow Cl^-_{(aq)}$$

The chlorine atoms are balanced as:

$$\overset{0}{Cl}_{2(s)} \longrightarrow \overset{-1}{Cl}^-_{(aq)}$$

The oxidation number is balanced by adding electrons

$$Cl_{2(s)} + 2e^- \longrightarrow 2Cl^-_{(aq)} \quad .....(ii)$$

The balanced chemical equation can be obtained by adding equation (i) and (ii) as:

$$Cl_{2(s)} + SO_{2(aq)} + 2H_2O_{(l)} \longrightarrow 2Cl^-_{(aq)} + SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)}$$

Page : 282 , Block Name : Exercise

Q8.24 Refer to the periodic table given in your book and now answer the following questions:
(a) Select the possible non metals that can show disproportionation reaction.
(b) Select three metals that can show disproportionation reaction.

Answer. In disproportionation reactions, one of the reacting substances always contains an element that can exist in at least three oxidation states,

(a) P, Cl, and S can show disproportionation reactions as these elements can exist in three or more oxidation states,
(b) Mn, Cu, and Ga can show disproportionation reactions as these elements can exist in three or more oxidation states.

Page : 282 , Block Name : Exercise

Q8.25 In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen?

Answer. The balanced chemical equation for the given reaction is given as:

$$\begin{array}{cccc} 4NH_{3(g)} + 5O_{2(g)} & \longrightarrow & 4NO_{(g)} & + 6H_2O_{(g)} \\ 4\times 17\,g & 5\times 32\,g & 4\times 30\,g & 6\times 18\,g \\ = 68\,g & = 160\,g & = 120\,g & = 108\,g \end{array}$$

Thus, 68 g of NH_3 reacts with 160 g of O_2
Therefore, 10g of NH3 reacts with $\frac{160\times 10}{68}$g of O_2, or 23.53 g of O_2.
But the available amount of $O_2$is 20 g.
Therefore, O_2 is the limiting reagent (we have considered the amount of O_2 to calculate the weight of nitric oxide obtained in the reaction).
Now, 160 g of O_2gives 120g of NO.
Therefore, 20 g of O_2gives$\frac{120\times 20}{160}$g of N, or 15 g of NO.
Hence, a maximum of 15 g of nitric oxide can be obtained.

Page : 282 , Block Name : Exercise

Q8.26 Using the standard electrode potentials given in the Table 8.1, predict if the reaction between the following is feasible:

(a) $Fe^{3+}(aq)$ and $I^{-}(aq)$

(b) $Ag^{+}(aq)$ and $Cu(s)$

(c) $Fe^{3+}(aq)$ and $Cu(s)$

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

(e) $Br_2(aq)$ and $Fe^{2+}(aq)$

Answer. (a) The possible reaction between $Fe^{3+}_{(aq)} + I^{-}_{(aq)}$ is given by

$$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$$

$$\text{Oxidation half equation : } \quad 2I^{-}_{(aq)} \longrightarrow I_{2(s)} + 2e^{-}; \quad E° = -0.54V$$

$$\text{Reduction half equation : } \quad [Fe^{3+}_{(aq)} + e^{-} \longrightarrow Fe^{2+}_{(aq)}] \times 2; \quad E° = +0.77V$$

$$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}; E° = +0.23V$$

E° for the overall reaction is positive. Thus, the reaction between $Fe^{3+}_{(aq)}$ and $I^{-}_{(aq)}$ is feasible.

(b) The possible reaction between $Ag^{+}_{(aq)} + Cu_{(s)}$ is given by,

$$2Ag^{+}_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$$

$$\text{Oxidation half equation : } \quad Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-} \quad ; E° = -0.34V$$

$$\text{Reduction half equation : } [Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)}] \times 2 \quad ; E° = +0.80V$$

$$2Ag^{+}_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}; E° = +0.46V$$

E° positive for the overall reaction is positive. Hence, the reaction between $Ag^{+}_{(aq)} + Cu_{(s)}$ is feasible.

(c) The possible reaction between $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$ is given by,

$$2Fe^{3+}_{(aq)} + Cu_{(s)} \longrightarrow 2Fe^{2+}_{(s)} + Cu^{2+}_{(aq)}$$

$$\text{Oxidation half equation : } \quad Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^{-} \quad ; E° = -0.34V$$

$$\text{Reduction half equation : } [Fe^{3+}_{(aq)} + e^{-} \longrightarrow Fe^{2+}_{(s)}] \times 2 \quad ; E° = +0.77V$$

$$2Fe^{3+}_{(aq)} + Cu_{(s)} \longrightarrow 2Fe^{2+}_{(s)} + Cu^{2+}_{(aq)}; E° = +0.43V$$

E° positive for the overall reaction is positive. Hence, the reaction between $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$ is feasible.

(d) The possible reaction between $Fe^{3+}_{(aq)}$ and $Ag_{(s)}$ is given by,

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

$$\text{Oxidation half equation : } \quad Ag_{(s)} \longrightarrow Ag^{+}_{(aq)} + e^{-} \quad ; E° = -0.80V$$

$$\text{Reduction half equation : } \quad Fe^{3+}_{(aq)} + e^{-} \longrightarrow Fe^{2+}_{(aq)} \quad ; E° = +0.77V$$

$$Ag_{(s)} + Fe^{3+}_{(aq)} \longrightarrow Ag^{+}_{(aq)} + Fe^{2+}_{(aq)}; E° = -0.03V$$

Here, E° for the overall reaction is negative. Hence, the reaction between $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$ is not feasible.

(e) The possible reaction between $Br_{2(aq)}$ and $Fe^{2+}_{(aq)}$ is given by,

$$Br_{2(s)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^{-}_{(aq)} + 2Fe^{3+}_{(aq)}$$

$$\begin{aligned}
&\text{Oxidation half equation :} && Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^{-}]\times 2 && ; E^\circ = -0.77V\\
&\text{Reduction half equation :} && Br_{2(aq)} + 2e^{-} \longrightarrow 2Br^{-}_{(aq)} && ; E^\circ = +1.09V\\
&\hline
&&& Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^{-}_{(aq)} + 2Fe^{3+}_{(aq)} && ; E^\circ = -0.32V
\end{aligned}$$

Here, E° for the overall reaction is positive. Hence, the reaction between $Br_{2(aq)}$ and $Fe^{2+}_{(aq)}$ is feasible.

Page : 282 , Block Name : Exercise

Q8.27 Predict the products of electrolysis in each of the following:
(i) An aqueous solution of $AgNO_3$ with silver electrodes
(ii) An aqueous solution $AgNO_3$ with platinum electrodes
(iii) A dilute solution of H_2SO_4 with platinum electrodes
(iv) An aqueous solution of $CuCl_2$ with platinum electrodes.
Answer. (i) $AgNO_3$ ionizes in aqueous solutions to form Ag^+ and NO_3^- ions.On electrolysis, either Ag^+ ions or H_2O molecules can be reduced at the cathode. But the reduction potential of Ag^+ ions is higher than that of H_2O.

$$Ag^{+}_{(aq)} + e^{-} \longrightarrow Ag_{(s)} ; E^\circ = +0.80V$$
$$2H_2O_{(l)} + 2e^{-} \longrightarrow H_{2(g)} + 2OH^{-}_{(aq)} ; E^\circ = -0.83V$$

Hence, Ag^+ ions are reduced at the cathode. Similarly, Ag metal or H_2O molecules can be oxidized at the anode. But the oxidation potential of Ag is higher than that of H_2O molecules.

$$Ag_{(s)} \longrightarrow Ag^{+}_{(aq)} + e^{-} \quad ; E^\circ = -0.80V$$
$$2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^{+}_{(aq)} + 4e^{-} ; E^\circ = -1.23V$$

Therefore, Ag metal gets oxidized at the anode.
(ii) Pt cannot be oxidized easily. Hence, at the anode, oxidation of water occurs to liberate O_2. At the cathode, Ag^+ ions are reduced and get deposited.
(iii) H_2SO_4 ionizes in aqueous solutions to give H+ and SO_4^{2-} ions.
$H_2SO_{4(aq)} \longrightarrow 2H^{+}_{(aq)} + SO_4^{2-}$
On electrolysis, either of H^+ ions or H_2O molecules can get reduced at the cathode. But the reduction potential of H^+ ions is higher than that of H_2Omolecules.

$$2H^{+}_{(aq)} + 2e^{-} \longrightarrow H_{2(g)} ; E^\circ = 0.0V$$
$$2H_2O_{(aq)} + 2e^{-} \longrightarrow H_{2(g)} + 2OH^{-}_{(aq)} ; E^\circ = -0.83V$$

Hence, at the cathode, H^+ ions are reduced to liberate H_2 gas.

On the other hand, at the anode, either of SO_4^{2-} ions or H_2O molecules can get oxidized. But the oxidation of SO_4^{2-} involves breaking of more bonds than that of H_2O molecules. Hence, SO_4^{2-} ions have a lower oxidation potential thanH_2O. Thus, H_2O is oxidized at the anode to liberate O_2 molecules.
(iv) In aqueous solutions, $CuCl_2$ ionizes to give Cu^{2+} and Cl^- ions as:
$CuCl_{2(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Cl^-_{(aq)}$
On electrolysis, either of Cu^{2+} ions or H_2O molecules can get reduced at the cathode. But the reduction potential of Cu^{2+} is more than that of H_2O molecules.

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(aq)} ; E° = +0.34V$$
$$H_2O_{(l)} + 2e^- \longrightarrow H_{2(g)} + 2OH^- ; E° = -0.83V$$

Hence,Cu^{2+} ions are reduced at the cathode and get deposited.
Similarly, at the anode, either of Cl^- or H_2O is oxidized. The oxidation potential of H_2O is higher than that of Cl^-.

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^- \quad ; E° = -1.36V$$
$$2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^- ; E° = -1.23V$$

But oxidation of H2O molecules occurs at a lower electrode potential than that of Cl^- ions because of over-voltage (extra voltage required to liberate gas). As a result, Cl^-.
ions are oxidized at the anode to liberate Cl_2 gas.
Similarly, at the anode, either of Cl– or H2O is oxidized. The oxidation potential of H2O is higher than that of Cl^-.

Page : 283 , Block Name : Exercise

Q8.28 Arrange the following metals in the order in which they displace each other from the solution of their salts.
Al, Cu, Fe, Mg and Zn.

Answer. A metal of stronger reducing power displaces another metal of weaker reducing power from its solution of salt.
The order of the increasing reducing power of the given metals is Cu < Fe < Zn < Al < Mg.
Hence, we can say that Mg can displace Al from its salt solution, but Al cannot displace Mg.
Thus, the order in which the given metals displace each other from the solution of their salts is given below:
Mg>Al> Zn> Fe,>Cu

Page : 283 , Block Name : Exercise

Q8.29 Given the standard electrode potentials,

$\mathrm{K^+/K = -2.93V, Ag^+/Ag = 0.80V}$
$\mathrm{Hg^{2+}/Hg = 0.79V}$
$\mathrm{Mg^{2+}/Mg = -2.37V \cdot Cr^{3+}/Cr = -0.74V}$
Arrange these metals in their increasing order of reducing power.

Answer. The lower the electrode potential, the stronger is the reducing agent. Therefore, the increasing order of the reducing power of the given metals is Ag < Hg < Cr < Mg < K.

Page : 283 , Block Name : Exercise

Q8.30 Depict the galvanic cell in which the reaction $\mathrm{Zn(s) + 2Ag^+(aq) \hat{a}t' Zn^{2+}(aq) + 2Ag(s)}$ takes place, further show:
(i) which of the electrode is negatively charged,
(ii) the carriers of the current in the cell, and
(iii) individual reaction at each electrode.

Answer. The galvanic cell corresponding to the given redox reaction can be represented as:
$\mathrm{Zn \left| Zn^{2+}_{(aq)} \right\| Ag^+_{(aq)} | Ag}$
(i) Zn electrode is negatively charged because at this electrode, Zn oxidizes to Zn2+ and the leaving electrons accumulate on this electrode.
(ii) Ions are the carriers of current in the cell.
(iii) The reaction taking place at Zn electrode can be represented as:
$\mathrm{Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-}$
And the reaction taking place at Ag electrode can be represented as:
$\mathrm{Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}}$

Page : 283 , Block Name : Exercise