NCERT Solutions for Class 11 Chemistry : Chapter 8 (Redox Reactions)

1.Assign oxidation number to the underlined elements in each of the following species:

(i) $NaH_2\underline{P}O_4$

(ii) $NaH\underline{S}O_4$

(iii) $H_4\underline{P}_2O_7$

(iv) $K_2\underline{Mn}O_4$

(v) $Ca\underline{O}_2$

(vi) $Na\underline{B}H_4$

(vii) $H_2\underline{S}_2O_7$

(viii) $KAl(\underline{S}O_4)_2.12H_2O$

Answer:

(i) $NaH_2\underline{P}O_4$

Let x be the oxidation no. of P.
Oxidation no. of Na = +1
Oxidation no. of H = +1
Oxidation no. of O = -2

$\overset{+1}{Na}\overset{+1}{H_2}\overset{x}{P}\overset{-2}{O_4}$

Then,
1(+1) + 2(+1) + 1(x) + 4(-2) = 0
1 + 2 + x -8 = 0
x = +5
Therefore, oxidation no. of P is +5.

(ii) $NaH\underline{S}O_4$

+1 +1 x -2

$\overset{+1}{Na}\overset{+1}{H}\overset{x}{S}\overset{-2}{O_4}$

Let x be the oxidation no. of S.

Oxidation no. of Na = +1

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

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

1 + 1 + x -8 = 0

x = +6

Therefore, oxidation no. of S is +6.

(iii) $H_4\underline{P}_2O_7$

$\overset{+1}{H_4}\overset{x}{P_2}\overset{-2}{O_7}$

Let x be the oxidation no. of P.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

4(+1) + 2(x) + 7(-2) = 0

4 + 2x − 14 = 0

2x = +10

x = +5

Therefore, oxidation no. of P is +5.

(iv) $K_2\underline{Mn}O_4$

$\overset{+1}{K_2}\overset{x}{Mn}\overset{-2}{O_4}$

Let x be the oxidation no. of Mn.

Oxidation no. of K = +1

Oxidation no. of O = -2

Then,

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

2 + x − 8 = 0

x = +6

Therefore, oxidation no. of Mn is +6.

(v) $Ca\underline{O}_2$

$\overset{+2}{Ca}\overset{x}{O_2}$

Let x be the oxidation no. of O.

Oxidation no. of Ca = +2

Then,

(+2) + 2(x) = 0

2 + 2x = 0

2x = -2

x = -1

Therefore, oxidation no. of O is -1.

(vi) $Na\underline{B}H_4$

$\overset{+1}{Na}\overset{x}{B}\overset{-1}{H_4}$

Let x be the oxidation no. of B.

Oxidation no. of Na = +1

Oxidation no. of H = -1

Then,

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

1 + x -4 = 0

x = +3

Therefore, oxidation no. of B is +3.

(vii) $H_2\underline{S}_2O_7$

$\overset{+1}{H_2}\overset{x}{S_2}\overset{-2}{O_7}$

Let x be the oxidation no. of S.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

2(+1) + 2(x) + 7(-2) = 0

2 + 2x − 14 = 0

2x = +12

x = +6

Therefore, oxidation no. of S is +6.

(viii) $KAl(\underline{S}O_4)_2.12H_2O$

$\overset{+1}{K}\overset{3+}{Al}\left(\overset{x}{S}\overset{2-}{O_4}\right)_2.12\overset{+1}{H_2}\overset{-2}{O}$

Let x be the oxidation no. of S.

Oxidation no. of K = +1

Oxidation no. of Al = +3

Oxidation no. of O= -2

Oxidation no. of H = +1

Then,

1(+1) + 1(+3) + 2(x) + 8(-2) + 24(+1) + 12(-2) = 0

1 + 3 + 2x − 16 + 24 − 24 = 0

2x = +12

x = +6

Therefore, oxidation no. of S is +6.

OR

Ignore the water molecules because it is neutral. Then, the summation of the oxidation no. of all atoms of water molecules can be taken as 0. Hence, ignore the water molecule.

1(+1) + 1(+3) + 2(x) + 8(-2) = 0

1 + 3 + 2x -16 = 0

2x = 12

x = +6

Therefore, oxidation no. of S is +6.

2.What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results ?

(i) $K\underline{I}_3$

(ii) $H_2\underline{S}_4O_6$

(iii) $\underline{Fe}_3O_4$

(iv) $\underline{C}H_3\underline{C}H_2OH$

(v) $\underline{C}H_3\underline{C}OOH$

Answer:

(i) $K\underline{I}_3$

Let x be the oxidation no. of I.

Oxidation no. of K = +1

Then,

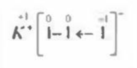
1(+1) + 3(x) = 0

1 + 3x = 0

x = $-\frac{1}{3}$

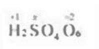
Oxidation no. cannot be fractional. Hence, consider the structure of KI_3.

In KI_3 molecule, an iodine atom forms coordinate covalent bond with an iodine molecule.



Therefore, in KI_3 molecule, the oxidation no. of I atoms forming the molecule I_2 is 0, while the oxidation no. of I atom which is forming coordinate bond is -1.

(ii) $H_2\underline{S_4}O_6$



Let x be the oxidation no. of S.

Oxidation no. of H = +1

Oxidation no. of O = -2

Then,

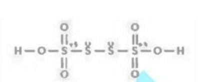
2(+1) + 4(x) + 6(-2) = 0

2 + 4x -12 = 0

4x = 10

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

Oxidation no. cannot be fractional. Therefore, S would be present with different oxidation state in molecule.



The oxidation no. of two out of the four S atoms is +5 while that of other two atoms is 0.

(iii) $\underline{Fe_3}O_4$

Let x be the oxidation no. of Fe.

Oxidation no. of O = -2

Then,

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

3x -8 = 0

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

Oxidation no. cannot be fractional.

One of the three atoms of Fe has oxidation no. +2 and other two atoms of Fe has oxidation no. +3.

$FeO.Fe_2O_3$

(iv) $\underline{C}H_3\underline{C}H_2OH$

C_2H_6O

Let x be the oxidation no. of C.

Oxidation no. of O= -2

Oxidation no. of H = +1

Then,

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

2x + 4 -2 = 0

x = -2

Therefore, oxidation no. of C is -2.

(v) $\underline{C}H_3\underline{C}OOH$

$C_2H_4O_2$

Let x be the oxidation no. of C.

Oxidation no. of O= -2

Oxidation no. of H = +1

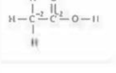
Then,

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

2x + 4 -4 = 0

x = 0

Therefore, average oxidation no. of C is 0. Both the carbon atoms are present in different environments so they cannot have same oxidation no. Therefore, carbon has oxidation no. of +2 and _2 in CH_3COOH .



3. Justify that the following reactions are redox reactions:

(i) $CuO_{(s)} + H_{2\,(g)} \rightarrow Cu_{(s)} + H_2O_{(g)}$

(ii) $Fe_2O_{3\,(s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(s)} + 3\ CO_{2\,(g)}$

(iii) $4\ BCl_{3\,(g)} + 3\ LiAlH_{4\,(s)} \rightarrow 2\ B_2H_{6\,(g)} + 3\ LiCl_{(s)} + 3\ AlCl_{3\,(s)}$

(iv) $2\ K_{(s)} + F_{2\ (g)} \rightarrow 2\ K + F_{(s)}$

(v) $4\ NH_{3\ (g)} + 5\ O_{2\ (g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$

Answer:

(i) $CuO_{(s)} + H_{2\ (g)} \rightarrow Cu_{(s)} + H_2O_{(g)}$

Oxidation no. of Cu and O in CuO is +2 and -2 respectively.

Oxidation no. of H_2 is 0.

Oxidation no. of Cu is 0.

Oxidation no. of H and O in H_2O is +1 and -2 respectively.

The oxidation no. of Cu decreased from +2 in CuO to 0 in Cu. That is CuO is reduced to Cu.

The oxidation no. of H increased from 0 to +1 in H_2. That is H_2 is oxidized to H_2O.

Therefore, the reaction is redox reaction.

(ii) $Fe_2O_{3\ (s)} + 3\ CO_{(g)} \rightarrow 2\ Fe_{(s)} + 3\ CO_{2\ (g)}$

In the above reaction,

Oxidation no. of Fe and O in Fe_2O_3 is +3 and -2 respectively.

Oxidation no. of C and O in CO is +2 and -2 respectively.
Oxidation no. of Fe is 0.

Oxidation no. of C and O in CO_2 is +4 and -2 respectively.

The oxidation no. of Fe decreased from +3 in Fe_2O_3 to 0 in Fe. That is Fe_2O_3 is reduced to Fe.

The oxidation no. of C increased from 0 to +2 in CO to +4 in CO_2. That is CO is oxidized to CO_2.

Therefore, the reaction is redox reaction.

(iii) $4\ BCl_{3\ (g)}\ +\ 3\ LiAlH_{4\ (s)}\ \rightarrow\ 2\ B_2H_{6\ (g)}\ +\ 3\ LiCl_{(s)}\ +\ 3\ AlCl_{3\ (s)}$

the above reaction,

Oxidation no. of B and Cl in BCl_3 is +3 and -1 respectively.

Oxidation no. of Li, Al and H in $LiAlH_4$ is +1, +3 and -1 respectively.

Oxidation no. of B and H in B_2H_6 is -3 and +1 respectively.

Oxidation no. of Li and Cl in LiCl is +1 and -1 respectively.

Oxidation no. of Al and Cl in $AlCl_3$ is +3 and -1 respectively.

The oxidation no. of B decreased from +3 in BCl_3 to -3 in B_2H_6. That is BCl_3 is reduced to B_2H_6.

The oxidation no. of H increased from -1 in $LiAlH_4$ to +1 in B_2H_6. That is $LiAlH_4$ is oxidized to

B_2H_6.

Therefore, the reaction is redox reaction.

(iv) $2\ K_{(s)}\ +\ F_{2\ (g)}\ \rightarrow\ 2\ K\ +\ F_{(s)}$

In the above reaction,
Oxidation no. of K is 0.
Oxidation no. of F is 0.
Oxidation no. of K and F in KF is +1 and -1 respectively.
The oxidation no. of K increased from 0 in K to +1 in KF. That is K is oxidized to KF.

The oxidation no. of F decreased from 0 in F_2 to -1 in KF. That is F_2 is reduced to KF.

Therefore, the reaction is a redox reaction.

(v) $4\ NH_{3\ (g)}\ +\ 5\ O_{2\ (g)}\ \rightarrow\ 4\ NO_{(g)}\ +\ 6\ H_2O_{(g)}$

In the above reaction,

Oxidation no. of N and H in NH_3 is -3 and +1 respectively.

Oxidation no. of O_2 is 0.

Oxidation no. of N and O in NO is +2 and -2 respectively.

Oxidation no. of H and O in H_2O is +1 and -2 respectively.

The oxidation no. of N increased from -3 in NH_3 to +2 in NO.

The oxidation no. of O_2 decreased from 0 in O_2 to -2 in NO and H_2O. That is O_2 is reduced.

Therefore, the reaction is a redox reaction.

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:

$$H_2O_{(s)} \ + \ F_{2\ (g)} \ \rightarrow \ HF_{(g)} \ + \ HOF_{(g)}$$

In the above reaction,

Oxidation no. of H and O in H_2O is +1 and -2 respectively.

Oxidation no. of F_2 is 0.

Oxidation no. of H and F in HF is +1 and -1 respectively.
Oxidation no. of H, O and F in HOF is +1, -2 and +1 respectively.

The oxidation no. of F increased from 0 in F_2 to +1 in HOF.

The oxidation no. of F decreased from 0 in O_2 to -1 in HF.

Therefore, F is both reduced as well as oxidized. So, it is redox reaction.

5. Calculate the oxidation no. of chromium, sulphur and nitrogen in H_2SO_5, $Cr_2O_7^{2-}$ and NO_3^-.

Give the structure for the compounds. Count for the fallacy.

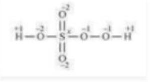
Answer:

(i) H_2SO_5

Let x be the oxidation no. of S.

Oxidation no. of O= -2

Oxidation no. of H = +1



Then,

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

2 + x − 10 = 0

x = +8

But the oxidation no. of S cannot be +8 as S has 6 valence electrons. Therefore, the oxidation no. of S cannot be more than +6.

The structure of H_2SO_5 is as given below:

Now,

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

2 + x − 6 − 2 = 0

x = +6

Therefore, the oxidation no. of S is +6.

(ii) $Cr_2O_7^{2-}$

Let x be the oxidation no. of Cr.

Oxidation no. of O= -2

Then,

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

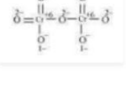
2x -14 = -2

x = +6

There is no fallacy about the oxidation no. of Cr in $Cr_2O_7^{2-}$.

The structure of $Cr_2O_7^{2-}$ is as given below.

Each of the two Cr atoms has the oxidation no. of +6.



(iii) NO_3^-

Let x be the oxidation no. of N.

Oxidation no. of O= -2

Then,

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

x − 6 = -1

x = +5

There is no fallacy about the oxidation no. of N in NO_3^-.

The structure of NO_3^- is as given below.

Nitrogen atom has the oxidation no. of +5.

6. Write formulas for the following compounds:

(i) Mercury

(II) chloride

(ii) Nickel (II) sulphate

(iii) Tin (IV) oxide

(iv) Thallium (I) sulphate

(v) Iron (III) sulphate

(vi) Chromium (III) oxide

Answer:

(i) Mercury (II) chloride

$HgCl_2$

(ii) Nickel (II) sulphate

$NiSO_4$

(iii) Tin (IV) oxide

SnO_2

(iv) Thallium (I) sulphate

Tl_2SO_4

(v) Iron (III) sulphate

$Fe_2(SO_4)_3$

(vi) Chromium (III) oxide

Cr_2O_3

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 compound where carbon has oxidation no. from -4 to +4 is as given below in the table:

Compounds	Oxidation no. of carbon
CH_2Cl_2	0
$HC \equiv CH$	-1
$ClC \equiv CCl$	+1
CH_3Cl	-2
$CHCl_3$, CO	+2
$H_3C - CH_3$	-3
$Cl_3C - CCl_3$	+3
CH_4	-4
CCl_4 , CO_2	+4

Compounds	Oxidation no. of nitrogen
N_2	0
N_2H_2	-1
N_2O	+1

N_2H_4	-2
NO	+2
NH_3	-3
N_2O_3	+3
NO_2	+4
N_2O_5	+5

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 (SO_2) the oxidation no. of S is +4 and the range of oxidation no. of sulphur is from +6 to -2.

Hence, SO_2 can act as reducing and oxidising agent.

In hydrogen peroxide (H_2O_2) the oxidation no. of O is -1 and the range of the oxidation no. of oxygen is from 0 to -2. Oxygen can sometimes attain the oxidation no. +1 and +2.

Therefore, H_2O_2 can act as reducing and oxidising agent.

In ozone (O_3) the oxidation no. of O is 0 and the range of the oxidation no. of oxygen is from 0 to −2. Hence, the oxidation no. of oxygen only decreases in this case.

Therefore, O_3 acts only as an oxidant.

In nitric acid (HNO_3) the oxidation no. of nitrogen is +5 and the range of the oxidation no. that nitrogen can have is from +5 to -3. Hence, the oxidation no. of nitrogen can only decrease in this case.

Therefore, HNO_3 acts only as an oxidant.

9.Consider the reactions:

(i) $6\ CO_{2\ (g)}\ +\ 6\ H_2O_{(l)}\ \rightarrow\ C_6H_{12}O_{6\ (aq)}\ +\ 6\ O_{2\ (g)}$

(ii) $O_{3\ (g)}\ +\ H_2O_{2\ (l)}\ \rightarrow\ H_2O_{(l)}\ +\ 2\ O_{2\ (g)}$

Why it is more appropriate to write these reactions as :

(i) $6\ CO_{2\ (g)}\ +\ 12\ H_2O_{(l)}\ \rightarrow\ C_6H_{12}O_{6\ (aq)}\ +\ 6\ H_2O_{(l)}\ +\ 6\ O_{2\ (g)}$

(ii) $O_{3\ (g)}\ +\ H_2O_{2\ (l)}\ \rightarrow\ H_2O_{(l)}\ +\ O_{2\ (g)}\ +\ O_{2\ (g)}$

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

(i)

Step 1 :

H_2O breaks to give H_2 and O_2.

$$2\ H_2O_{(l)}\ \rightarrow\ 2\ H_{2\ (g)}\ +\ O_{2\ (g)}$$

Step 2 :

The H_2 produced in earlier step reduces CO_2, thus produce glucose and water.

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

The net reaction is as given below:

$[2\ H_2O_{(l)}\ \rightarrow\ 2\ H_{2\ (g)}\ +\ O_{2\ (g)}] \times 6$

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

$$6\ CO_{2\ (g)}\ +\ 12\ H_2O_{(l)}\ \rightarrow\ C_6H_{12}O_{6\ (g)}\ +\ 6\ H_2O_{(l)}\ +\ 6\ O_{2\ (g)}$$

This is the suitable way to write the reaction as the reaction also produce water molecules in the photosynthesis process.

The path can be found with the help of radioactive H_2O^{18} instead of H_2O.

(ii)

Step 1

O_2 is produced from each of the reactants O_3 and H_2O_2. That is the reason O_2 is written two times.

O_3 breaks to form O_2 and O.

Step 2 :

H_2O_2 reacts with O produced in the earlier step, thus produce H_2O and O_2.

$$O_{3\,(g)} \rightarrow O_{2\,(g)} + O_{(g)}$$

$$H_2O_{2\,(l)} + O_{(g)} \rightarrow H_2O_{(l)} + O_{2\,(g)}$$

$$H_2O_{2\,(l)} + O_{3\,(g)} \rightarrow H_2O_{(l)} + O_{2\,(g)} + O_{2\,(g)}$$

The path can be found with the help of $H_2O_2^{18}$ or O_3^{18}.

10. The compound AgF2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why ?

Answer:

The oxidation no. of Ag in AgF_2 is +2. But, +2 is very unstable oxidation no. of Ag. Hence, when AgF_2 is formed, silver accepts an electron and forms Ag^+. This decreases the oxidation no. of Ag from +2 to +1. +1 state is more stable. Therefore, AgF_2 acts as a very strong oxidizing agent.

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.

Justify the above statement with three examples.

Answer:

When there is a reaction between reducing agent and oxidizing agent, a compound is formed which has lower oxidation number if the reducing agent is in excess and a compound is formed which has higher oxidation number if the oxidizing agent is in excess.

(i) P_4 and F_2 are reducing and oxidizing agent respectively.

In an excess amount of P_4 is reacted with F_2, then PF_3 would be produced, where the oxidation no. of P is +3.

$$P_{4\ (excess)}\ F_2 \rightarrow PF_3$$

If P_4 is reacted with excess of F_2, then PF_5 would be produced, where the oxidation no. of P is +5.

$$P_4 + F_{2\ (excess)} \rightarrow PF_5$$

(ii) K and O_2 acts as a reducing agent and oxidizing agent respectively.

If an excess of K reacts with O_2, it produces K_2O. Here, the oxidation number of O is -2.

$$4\ K_{\ (excess)} + O_2 \rightarrow 2\ K_2O^{-2}$$

If K reacts with an excess of O_2, it produces K_2O_2, where the oxidation number of O is −1.

$$2\ K + O_{2\ (excess)} \rightarrow K_2O_2^{-1}$$

(iii) C and O_2 acts as a reducing agent and oxidizing agent respectively.

If an excess amount of C is reacted with insufficient amount of O_2, then it produces CO, where the oxidation number of C is +2.

$$C_{\ (excess)} + O_2 \rightarrow CO$$

If C is burnt in excess amount of O_2, then CO_2 is produced, where the oxidation number of C is +4.

$$C + O_{2\ (excess)} \rightarrow CO_2$$

12.How do you count for the following observations?

(i) Acidic potassium permanganate and alkaline 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.

(ii) 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?

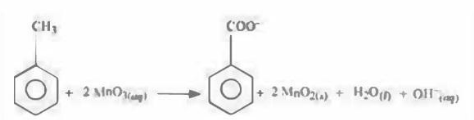
Answer:

(i) While manufacturing benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant due to the given reasons.

(a) In a neutral medium, OH^- ions are produced in the reaction. Due to that, the cost of adding an acid or a base can be reduced.

(b) $KMnO_4$ and alcohol are homogeneous to each other as they are polar. Alcohol and toluene are homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium compared to heterogeneous medium. Therefore, in alcohol, $KMnO_4$ and toluene can react at a faster rate.

The redox reaction is as given below:



(ii) When concentrated H_2SO_4 is added to an inorganic mixture containing bromide, firstly HBr is produced. HBr, a strong reducing agent, reduces H_2SO_4 to SO_2 with the evolution of bromine's red vapour.

$$2\,NaBr \;+\; 2\,H_2SO_4 \;\rightarrow\; 2\,NaHSO_4 \;+\; 2\,HBr$$

$$2\,HBr \;+\; H_2SO_4 \;\rightarrow\; Br_2 \;+\; SO_2 \;+\; 2\,H_2O$$

When concentrated H_2SO_4 I added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, a weak reducing agent, cannot reduce H_2SO_4 to SO_2.

$$2\,NaCl \;+\; 2\,H_2SO_4 \;\rightarrow\; 2\,NaHSO_4 \;+\; 2\,HCl$$

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

(i) $2\ AgBr_{(s)}\ +\ C_6H_6O_{2\ (aq)}\ \rightarrow\ 2\ Ag_{(s)}\ +\ 2\ HBr_{(aq)}\ +\ C_6H_4O_{2\ (aq)}$

(ii) $HCHO_{(l)}\ +\ 2\ [Ag(NH_3)_2]^+_{(aq)}\ +\ 3\ OH^-_{(aq)}\ \rightarrow\ 2\ Ag_{(s)}\ +\ HCOO^-_{(aq)}\ +$

$4\ NH_{3\ (aq)}\ +\ 2\ H_2O_{(l)}$

(iii) $HCHO_{(l)}\ +\ 2\ Cu^{2+}_{(aq)}\ +\ 5\ OH^-_{(aq)}\ \rightarrow\ Cu_2O_{(s)}\ +\ HCOO^-_{(aq)}\ +\ 3\ H_2O_{(l)}$

(iv) $N_2H_{4\ (l)}\ +\ 2\ H_2O_{2\ (l)}\ \rightarrow\ N_{2\ (g)}\ +\ 4\ H_2O_{(l)}$

(v) $Pb_{(s)}\ +\ PbO_{2\ (s)}\ +\ 2\ H_2SO_{4\ (aq)}\ \rightarrow\ 2\ PbSO_{4\ (aq)}\ +\ 2\ H_2O_{(l)}$

Answer:

(i) $2\ AgBr_{(s)}\ +\ C_6H_6O_{2\ (aq)}\ \rightarrow\ 2\ Ag_{(s)}\ +\ 2\ HBr_{(aq)}\ +\ C_6H_4O_{2\ (aq)}$

$C_6H_6O_2$ => Oxidized substance

AgBr => Reduced substance

AgBr =>Oxidizing agent

$C_6H_6O_2$ => Reducing agent

(ii) $HCHO_{(l)}\ +\ 2\ [Ag(NH_3)_2]^+_{(aq)}\ +\ 3\ OH^-_{(aq)}\ \rightarrow\ 2\ Ag_{(s)}\ +\ HCOO^-_{(aq)}\ +$

$4\ NH_{3\ (aq)}\ +\ 2\ H_2O_{(l)}$

HCHO => Oxidized substance

$[Ag(NH_3)_2]^+$ => Reduced substance

$[Ag(NH_3)_2]^+$ => Oxidizing agent

HCHO=> Reducing agent

(iii) $HCHO_{(l)}\ +\ 2\ Cu^{2+}_{(aq)}\ +\ 5\ OH^-_{(aq)}\ \rightarrow\ Cu_2O_{(s)}\ +\ HCOO^-_{(aq)}\ +\ 3\ H_2O_{(l)}$

HCHO => Oxidized substance

Cu^{2+} => Reduced substance

Cu^{2+} => Oxidizing agent

HCHO => Reducing agent

(iv) $N_2H_{4\ (l)}\ +\ 2\ H_2O_{2\ (l)}\ \rightarrow\ N_{2\ (g)}\ +\ 4\ H_2O_{(l)}$

N_2H_4 => Oxidized substance

H_2O_2 => Reduced substance

H_2O_2 => Oxidizing agent

N_2H_4 => Reducing agent

(v) $Pb_{(s)}\ +\ PbO_{2\ (s)}\ +\ 2\ H_2SO_{4\ (aq)}\ \rightarrow\ 2\ PbSO_{4\ (aq)}\ +\ 2\ H_2O_{(l)}$

Pb=> Oxidized substance

PbO_2 => Reduced substance

PbO_2 => Oxidizing agent

Pb => Reducing agent

14. Consider the reactions :

$2\ S_2O^{2-}_{3\ (aq)}\ +\ I_{2\ (s)}\ \rightarrow\ S_4O^{2-}_{6\ (aq)}\ +\ 2\ I^-_{(aq)}$

$S_2O^{2-}_{3\ (aq)}\ +\ 2\ Br_{2\ (l)}\ +\ 5\ H_2O_{(l)}\ \rightarrow\ 2\ SO^{2-}_{4\ (aq)}\ +\ 4\ Br^-_{(aq)}\ +\ 10\ H^+_{(aq)}$

Why does the same reductant, thiosulphate react differently with iodine and bromine ?

Answer:

The average oxidation no. of S in $S_2O_3^{2-}$ is +2.

The average oxidation no. of S in $S_4O_6^{2-}$ is +2.5.

The oxidation no. of S in $S_2O_3^{2-}$ is +2.

The oxidation no. of S in SO_4^{2-} is +6.

As Br_2 is a stronger oxidizing agent than I_2, it oxidizes S of $S_2O_3^{2-}$ to a higher oxidation no. of +6 in SO_4^{2-}.

As I_2 is a weaker oxidizing agent so it oxidizes S of $S_2O_3^{2-}$ ion to a lower oxidation no. that is 2.5 in $S_4O_6^{2-}$ ions.

Thus, thiosulphate react differently with I_2 and Br_2.

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 as:

$$F_{2\,(aq)} + 2\,Cl^-_{(s)} \rightarrow 2\,F^-_{(aq)} + Cl_{2\,(g)}$$

$$F_{2\,(aq)} + 2\,Br^-_{(aq)} \rightarrow 2\,F^-_{(aq)} + Br_{2\,(l)}$$

$$F_{2\,(aq)} + 2\,I^-_{(aq)} \rightarrow 2\,F^-_{(aq)} + I_{2\,(s)}$$

But, Cl_2, Br_2, and I_2 cannot oxidize F^- to F_2. The oxidizing power of halogens increases in the order as given below:

$I_2 < Br_2 < Cl_2 < F_2$

Therefore, fluorine is the best oxidant among halogens.

HI and HBr can reduce H_2SO_4 to SO_2, but HCl and HF cannot. Hence, HI and HBr are stronger reductants compared to HCl and HF.

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

$$2\,HBr\;+\;H_2SO_4\;\rightarrow\;Br_2\;+\;SO_2\;+\;2\,H_2O$$

I^- can reduce Cu^{2+} to Cu^+, but Br^- cannot.

$$4\,I^-_{(aq)}\;+\;2\,Cu^{2+}_{(aq)}\;\rightarrow\;Cu_2I_{2\,(s)}\;+\;I_{2\,(aq)}$$

Therefore, hydrochloric acid is the best reductant among hydrohalic compounds.
Hence, the reducing power of hydrohalic acids increases as given below:

$$HF < HCl < HBr < HI$$

16. Why does the following reaction occur ?

$$XeO^{4-}_{6\,(aq)}\;+\;2\,F^-_{(aq)}\;+\;6\,H^+_{(aq)}\;\rightarrow\;XeO_{3\,(g)}\;+\;F_{2\,(g)}\;+\;3\,H_2O_{(l)}$$

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

Answer:

$$XeO^{4-}_{6\,(aq)}\;+\;2\,F^-_{(aq)}\;+\;6\,H^+_{(aq)}\;\rightarrow\;XeO_{3\,(g)}\;+\;F_{2\,(g)}\;+\;3\,H_2O_{(l)}$$

The oxidation no. of Xe reduces from +8 in XeO_6^{4-} to +6 in XeO_3.

The oxidation no. of F increases from -1 in F^- to 0 in F_2.

Hence, XeO_6^{4-} is reduced on the other hand F^- is oxidized. As $Na_2XeO_6^{4-}$ (or XeO_6^{4-}) is a stronger oxidizing agent compared to F_2, this reaction occurs.

17. Consider the reactions:

(i) $H_3PO_{2\,(aq)}\;+\;4\,AgNO_{3\,(aq)}\;+\;2\,H_2O_{(l)}\;\rightarrow\;H_3PO_{4\,(aq)}\;+\;4\,Ag_{(s)}\;+\;4\,HNO_{3\,(aq)}$

(ii) $H_3PO_{2\,(aq)}\;+\;2\,CuSO_{4\,(aq)}\;+\;2\,H_2O_{(l)}\;\rightarrow\;H_3PO_{4\,(aq)}\;+\;2\,Cu_{(s)}\;+\;H_2SO_{4\,(aq)}$

(iii) $C_6H_5CHO_{(l)} \; + \; 2\,[Ag(NH_3)_2]^+_{(aq)} \; + \; 3\,OH^-_{(aq)} \; \rightarrow \; C_6H_5COO^-_{(aq)} \; +$

$2\,Ag_{(s)} \; + \; 4\,NH_{3\,(aq)} \; + \; 2\,H_2O_{(l)}$

(iv) $C_6H_5CHO_{(l)} \; + \; 2\,Cu^{2+}_{(aq)} \; + \; 5\,OH^-_{(aq)} \; \rightarrow$ ***No change is observed***

What can you inference from the reactions about the behavior of Ag^+ and Cu^{2+} ?

Answer:

Ag^+ and Cu^{2+} behaves as oxidizing agent in reactions (i) and (ii) respectively.

In reaction (iii), Ag^+ oxidizes C_6H_5CHO to $C_6H_5COO^-$

In reaction (iv), Cu^{2+} cannot oxidize C_6H_5CHO.

Therefore, Ag^+ is a stronger oxidizing agent compared to Cu^{2+}.

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

(i) $MnO^-_{4\,(aq)} \; + \; I^-_{(aq)} \; \rightarrow \; MnO_{2\,(s)} \; + \; I_{2\,(s)}$ ***(Basic medium)***

(ii) $MnO^-_{4\,(aq)} \; + \; SO_{2\,(g)} \; \rightarrow \; Mn^{2+}_{(aq)} \; + \; H_2SO^-_4$ ***(Acidic medium)***

(iii) $H_2O_{2\,(aq)} \; + \; Fe^{2+}_{(aq)} \; \rightarrow \; Fe^{3+}_{(aq)} \; + \; H_2O_{(l)}$ ***(Acidic medium)***

(iv) $Cr^{2-}_2O_{7\,(aq)} \; + \; SO_{2\,(g)} \; \rightarrow \; Cr^{3+}_{(aq)} \; + \; SO^{2-}_{(aq)}$ ***(Acidic medium)***

Answer:

(i) $MnO^-_{4\,(aq)} \; + \; I^-_{(aq)} \; \rightarrow \; MnO_{2\,(s)} \; + \; I_{2\,(s)}$

Step 1
The two half reactions are given below:

Oxidation half reaction: $I_{(aq)} \; \rightarrow \; I_{2\,(s)}$

Reduction half reaction: $MnO^-_4 \; \rightarrow \; MnO_2$

Step 2

Balance I in oxidation half reaction:

$$2\ I^-_{(aq)} \ \rightarrow \ I_{2\ (s)}$$

Add 2 e^- to the right hand side of the reaction to balance the charge:

$$2I^-_{(aq)} \ \rightarrow \ I_{2\ (s)} \ + \ 2\ e^-$$

Step 3

The oxidation no. of Mn has decreased from +7 to +4 in the reduction half reaction. Therefore, 3 electrons are added to the left hand side of the reaction.

$$MnO^-_{4\ (aq)} \ + \ 3\ e^- \ \rightarrow \ MnO_{2\ (aq)}$$

Add 4 OH^- ions to right hand side of the reaction to balance the charge.

$$MnO^-_{4\ (aq)} \ + \ 3\ e^- \ \rightarrow \ MnO_{2\ (aq)} \ + \ 4\ OH^-$$

Step 4

There are 6 oxygen atoms on the right hand side and 4 oxygen atoms on the left hand side. Hence, 2 water molecules are added to the left hand side.

$$MnO^-_{4\ (aq)} \ + \ 2\ H_2O \ + \ 3\ e^- \ \rightarrow \ MnO_{2\ (aq)} \ + \ 4OH^-$$

Step 5

Equal the no. of electrons on both the sides by multiplying oxidation half reaction by 3 and reduction half reaction by 2:

$$6\ I^-_{(aq)} \ \rightarrow \ 3\ I_{2\ (s)} \ + \ 6\ e^-$$

$$2\ MnO^-_{4\ (aq)} \ + \ 4\ H_2O \ + \ 6\ e^- \ \rightarrow 2\ MnO_{2\ (s)} \ + \ 8\ OH^-_{(aq)}$$

Step 6

After adding both the half reactions, we get the balanced reaction as given below:

$$6\ I^-_{(aq)} \ + \ 2\ MnO^-_{4\ (aq)} \ + \ 4\ H_2O_{(l)} \ \rightarrow \ 3\ I_{2\ (s)} \ + \ 2\ MnO_{2\ (s)} \ + \ 8\ OH^-_{(aq)}$$

(ii) $MnO^-_{4\ (aq)} \ + \ SO_{2\ (g)} \ \rightarrow \ Mn^{2+}_{(aq)} \ + \ H_2SO^-_4$

Step 1

Similar to (i), oxidation half reaction is:

$$SO_{2\,(g)} \ + \ 2\,H_2O_{(l)} \ \rightarrow \ HSO^-_{4\,(aq)} \ + \ 3\,H^+_{(aq)} \ + \ 2\,e^-_{(aq)}$$

Step 2

Reduction half reaction is:

$$MnO^-_{4\,(aq)} \ + \ 8\,H^+_{(aq)} \ + \ 5\,e^- \ \rightarrow \ Mn^{2+}_{(aq)} \ + \ 4\,H_2O_{(l)}$$

Step 3

Multiply the oxidation half reaction with 5 and the reduction half reaction with 2, then add them. We get the balanced reaction as given below:

$$2\,MnO^-_{4\,(aq)} \ + \ 5\,SO_{2\,(g)} \ + \ 2\,H_2O_{(l)} \ + \ H^+_{(aq)} \ \rightarrow \ 2\,Mn^{2+}_{(aq)} \ + \ 5\,HSO^-_{4\,(aq)}$$

(iii) $H_2O_{2\,(aq)} \ + \ Fe^{2+}_{(aq)} \ \rightarrow \ Fe^{3+}_{(aq)} \ + \ H_2O_{(l)}$

Step 1

Similar to (i), oxidation half reaction is:

$$Fe^{2+}_{(aq)} \ \rightarrow \ Fe^{3+}_{(aq)} \ + \ e^-$$

Step 2

Reduction half reaction is:

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

Step 3

Multiply the oxidation half reaction with 2 then add it to the reduction half reaction. We get the balanced reaction as given below:

$$H_2O_{2\,(aq)} \ + \ 2\,Fe^{2+}_{(aq)}\ 2\,H^+_{(aq)} \ \rightarrow \ 2\,Fe^{3+}_{(aq)} 2\,H_2O_{(l)}$$

(iv) $Cr^{2-}_2O_{7\,(aq)} \ + \ SO_{2\,(g)} \ \rightarrow \ Cr^{3+}_{(aq)} \ + \ SO^{2-}_{(aq)}$

Step 1

Similar to (i), oxidation half reaction is:

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

Step 2

Reduction half reaction is:

$$Cr_2O^{2-}_{7\,(aq)} \ + \ 14\,H^+_{(aq)} \ + \ 6\,e^- \ \rightarrow \ 2\,Cr^{3+}_{(aq)} \ + \ 7\,H_2O_{(l)}$$

Step 3

Multiply the oxidation half reaction with 2 then add it to the reduction half reaction. We get the balanced reaction as given below:

$$Cr_2^{2-}O_{7\ (aq)} \ + \ 3\ SO_{2\ (g)} \ + \ 2\ H^+_{(aq)} \ \rightarrow \ 2\ Cr^{3+}_{(aq)} \ + \ 3\ SO^{2-}_{4\ (aq)} \ + \ H_2O_{(l)}$$

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.

(i) $P_{4\ (s)} \ + \ OH^-_{(aq)} \ \rightarrow \ PH_{3\ (g)} \ + \ HPO^-_{2\ (aq)}$

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

(iii) $Cl_2O_{7\ (g)} \ + \ H_2O_{2\ (aq)} \ \rightarrow \ ClO^-_{2\ (aq)} \ + \ O_{2\ (g)} \ + \ H^+_{(aq)}$

Answer:

(i) The Oxidation no. of P reduces from 0 in P_4 to – 3 in PH_3

The oxidation no. of P increases from 0 in P_4 to + 2 in HPO_2^-. Therefore, P_4 behaves both as

areducing agent as well asoxidizing agent in the reaction.

Ion – electron method:

– The oxidation half reaction:

$$P_{4\ (s)} \ \rightarrow \ HPO^-_{2\ (aq)}$$

– Balance atom P:

$$P_{4\ (s)} \ \rightarrow \ 4\ HPO^-_{2\ (aq)}$$

– Add 8 electrons to balance oxidation no.

$$P_{4\ (s)} \ \rightarrow \ 4\ HPO^-_{2\ (aq)} \ + \ 8\ e^-$$

– Add $12\ OH^-$ to balance the charge:

$$P_{4\ (s)} \ + \ 12\ OH^-_{(aq)} \ \rightarrow \ 4\ HPO^-_{2\ (aq)} \ + \ 8\ e^-$$

– Add 4 H_2O to balance H and O atoms:

$$P_{4\,(s)} + 12\,OH^-_{(aq)} \rightarrow 4\,HPO^-_{2\,(aq)} + 4\,H_2O_{(l)} + 8\,e^- \text{ ----(1)}$$

– The reduction half reaction:

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

– Balance atom P:

$$P^0_{4\,(s)} \rightarrow 4\,P^{-3}H_{3\,(g)}$$

– Add 12 electrons to balance oxidation no.

$$P_{4\,(s)} + 12\,e^- \rightarrow 4\,PH_{3\,(g)}$$

– Add $12\ OH^-$ to balance the charge:

$$P_{4\,(s)} + 12\,e^- \rightarrow 4\,PH_{3\,(g)} + 12\,OH^-_{(aq)}$$

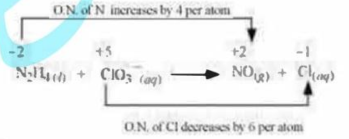
– Add 12 H_2O to balance H and O atoms:

$$P_{4\,(s)} + 12\,H_2O_{(l)} + 12\,e^- \rightarrow 4\,PH_{3\,(g)} + 12\,OH^-_{(aq)} \text{ ----(2)}$$

– Now, multiply the equation (1) by 3 and equation (2) by 2. Then, after adding them, we get the balanced redox reaction as given below:

$$5\,P_{4\,(s)} + 12\,H_2O_{(l)} + 12\,HO^-_{(aq)} \rightarrow 8\,PH_{3\,(g)} + 12\,HPO^-_{2\,(aq)}$$

(ii)



The Oxidation no. of N increases from -2 in N_2H_4 to −+2 in NO.

The oxidation no. of Cl reduces from +5 in ClO_3^- to +-1 in Cl^-.

Therefore, N_2H_4 behaves as a reducing agent while ClO_3^- behaves as an oxidizing agent in the reaction.

Ion – electron method:

– The oxidation half reaction:

$$N_2H_{4\,(l)} \rightarrow NO_{(g)}$$

– Balance atom N:

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

– Add 8 electrons to balance oxidation no:

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

– Add $8\,OH^-$ to balance the charge:

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

– Add 6 H_2O to balance O atoms:

$$N_2H_{4\,(l)} + 8\,OH^-_{(aq)} \rightarrow 2\,NO_{(g)} + 6\,H_2O_{(l)} + 8\,e^- \text{ --- (1)}$$

– The reduction half reaction:

$$ClO^-_{3\,(aq)} \rightarrow Cl^-_{(aq)}$$

– Add 6 electrons to balance oxidation no.

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

– Add $6\,OH^-$ ions to balance the charge:

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

– Add 3 H_2O to balance O atoms:

$$ClO^-_{3\,(aq)} + 3\,H_2O_{(l)} + 6\,e^- \rightarrow Cl^-_{(aq)} + 6\,OH^-_{(aq)} \text{ --- (2)}$$

Now, multiply the equation (1) by 3 and equation (2) by 4. Then, after adding them, we get the balanced redox reaction as given below:

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

Oxidation number method:

– Reduction in the oxidation no. of N = 2 × 4 = 8

– Increment in the oxidation no. of Cl = 1 × 6 = 6

Multiply N_2H_4 by 3 and ClO_3^- by 4 to balance the reduction and increment of the oxidation no. :

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

– Balance Cl and n atoms:

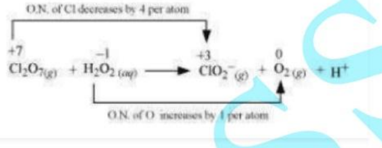
$$3\ N_2H_{4\ (l)}\ +\ 4\ ClO^-_{3\ (aq)}\ \rightarrow\ 6\ NO_{(g)}\ +\ 4\ Cl^-_{(aq)}$$

– Add 6 H_2O to balance O atoms:

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

This is the required reaction equation.

(iii)



The Oxidation no. of Cldecreases from +7 in Cl_2O_7 to +3 in ClO_2^-.

The oxidation no. of Oincreases from -1 in H_2O_2 to 0 in O_2.

Therefore, H_2O_2 behaves as a reducing agent while Cl_2O_7 behaves as an oxidizing agent in the reaction.

Ion – electron method:

– The oxidation half reaction:

$$H_2O_{2\ (aq)}\ \rightarrow\ O_{2\ (g)}$$

– Add 2 electrons to balance oxidation no:

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

– Add $2\ OH^-$ to balance the charge:

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

– Add 2 H_2O to balance O atoms:

$$H_2O_{2\ (aq)} \ + \ 2\ OH^-_{(aq)} \rightarrow \ O_{2\ (g)} \ + \ 2\ H_2O_{(l)} \ + \ 2\ e^-$$ --- (1)

– The reduction half reaction:

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

– Balance Cl atoms:

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

– Add 8 electrons to balance oxidation no.

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

– Add $6\ OH^-$ ions to balance the charge:

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

– Add 3 H_2O to balance O atoms:

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

Now, multiply the equation (1) by 4. Then, adding equation (1) and (2), we get the balanced redox reaction as given below:

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

Oxidation number method:

– Reduction in the oxidation no. of Cl_2O_7 = 4× 2 = 8

– Increment in the oxidation no. of H_2O_2 = 2× 1 = 2

Multiply H_2O_2 by 4 and O_2 by 4 to balance the reduction and increment of the oxidation no. :

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

– Balance Cl and n atoms:

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

– Add 3 H_2O to balance O atoms:

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

– Add $2\,OH^-$ and $2\,H_2O$ to balance H atoms:

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

This is the required reaction equation.

20.What sorts of informations can you draw from the following reaction ?

$$(CN)_{2\,(g)} \;+\; 2\,OH^-_{(aq)} \;\rightarrow\; CN^-_{(aq)} \;+\; CNO^-_{(aq)} \;+\; H_2O_{(l)}$$

Answer:

The oxidation no. of C in $(CN)_2$, CN^- and CNO^- are +3, +2 and +4 respectively.

Let the oxidation no. of C be y.

$(CN)_2$

2(y – 3) = 0

Therefore, y = 3

CN^-

y – 3 = -1

Therefore, y = 2

CNO^-

y – 3 -2 = -1

Therefore, y = 4

The oxidation no. of C in the reaction is:

Oxidation no. of C in $(CN)_2$ is +3

Oxidation no. of C in CN^- is +2

Oxidation no. of C in CNO^- is +4

We can see that the same compound is oxidized and reduced simultaneously in the reaction.

The reactions in which the same compound is oxidized and reduced is known as disproportionation reaction. Then, we can say that alkaline decomposition of cyanogens is a disproportionation reaction.

21. Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn2+, MnO2, and H+ ion. Write a balanced ionic equation for the reaction.

Answer:

The reaction is as given below:

$$Mn^{3+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + MnO_{2\,(s)} + H^{+}_{(aq)}$$

The oxidation half reaction:

$$Mn^{3+}_{(aq)} \rightarrow MnO_{2\,(s)}$$

Add 1 electron to balance the oxidation no. :

$$Mn^{3+}_{(aq)} \rightarrow MnO_{2\,(s)} + e^{-}$$

Add $4\ H^{+}$ ions to balance the charge:

$$Mn^{3+}_{(aq)} \rightarrow MnO_{2\,(s)} + e^{-} + 4\ H^{+}_{(aq)}$$

Add 2 H_2O to balance O atoms and H^{+} ions:

$$Mn^{3+}_{(aq)} + 2\ H_2O_{(l)} \rightarrow MnO_{2\,(s)} + e^{-} + 4\ H^{+}_{(aq)} \text{ ---- (1)}$$

The reduction half reaction:

$$Mn^{3+}_{(aq)} \rightarrow Mn^{2+}_{(aq)}$$

Add 1 electron to balance the oxidation no. :

$$Mn^{3+}_{(aq)} + e^{-} \rightarrow Mn^{2+}_{(aq)} \text{ --- (2)}$$

Add equation (1) and (2) to get the balanced chemical equation:

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

22.Consider the elements:

Cs, I, Ne and F

(i) Which element exhibit only negative oxidation no.?

(ii) Which element exhibit only positive oxidation no.?

(iii) Which element exhibit both negative and positive oxidation no.?

(iv) Which element exhibits neither negative nor positive oxidation no.?

Answer:

(i) F exhibits only negative oxidation no. That is -1.

(ii) Cs exhibits only positive oxidation no. That is +1.

(iii) I exhibits both negative and positive oxidation no. That is -1, +1, +3, +5 and +7.

(iv) Ne exhibits neither negative nor positive oxidation no. That is 0.

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 redox reaction is as given below:

$$Cl_{2\ (s)} \ + \ SO_{2\ (aq)} \ + \ H_2O_{(l)} \ \rightarrow \ Cl^-_{(aq)} \ + \ SO^{2-}_{4\ (aq)}$$

The oxidation half reaction:

$$SO_{2\ (aq)} \ \rightarrow \ SO^{2-}_{4\ (aq)}$$

Add 2 electrons to balance the oxidation no. :

$$SO_{2\ (aq)} \ \rightarrow \ SO^{2-}_{4\ (aq)} \ + \ 2\ e^-$$

Add $4\ H^+$ ions to balance the charge:

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

Add 2 H_2O to balance O atoms and H^+ ions:

$$SO_{2\ (aq)} \ + \ 2\ H_2O \ \rightarrow \ SO^{2-}_{4\ (aq)} \ + \ 4\ H^+_{(aq)} \ + \ 2\ e^- \ --- (1)$$

The reduction half reaction:

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

Balance Cl atoms:

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

Add 2 electrons to balance the oxidation no. :

$$Cl_{2\,(s)} \;+\; 2\,e^{-} \rightarrow \; 2\,Cl^{-}_{(aq)}$$ --- (2)

Add equation (1) and (2) to get the balanced chemical equation:

$$Cl_{2\,(s)} \;+\; SO_{2\,(aq)} \;+\; 2\,H_2O_{(l)} \;\rightarrow\; 2\,Cl^{-}_{(aq)} \;+\; SO^{2-}_{4\,(aq)} \;+\; 4\,H^{+}_{(aq)}$$

24. Refer to the periodic table given in your book and now answer the following questions:

(i) Which non – metals can show disproportionation reaction?

(ii) Which three metals shows disproportionation reaction?

Answer:

One of the reacting elements always has an element that can exist in at least 3 oxidation numbers.

(i) The non – metals which can show disproportionation reactions are P, Cl and S.

(ii) The three metals which can show disproportionation reactions are Mn, Ga and Cu.

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 reaction is as given below:

$$4\ NH_{3\ (g)}\ +\ 5\ O_{2\ (g)}\ \rightarrow\ 4\ NO_{(g)}\ +\ 6\ H_2O_{(g)}$$

$4\ NH_3$ = 4 × 17 g = 68 g

$5\ O_2$ = 5 × 32 g = 160 g

$4\ NO$ = 4 × 30 g = 120 g

$6\ H_2O$ = 6 × 18 g = 108 g

Thus, NH_3 (68 g) reacts with O_2 (20 g)

Therefore, 10 g of NH_3 reacts with $\frac{160\times10}{68}$ g = 23.53 g of O_2

But only 20 g of O_2 is available.

Hence, O_2 is a limiting reagent.

Now, 160 g of O_2 gives $\frac{120\times20}{160}$ g of N = 15 g of NO.

Therefore, max of 15 g of nitric oxide can be obtained.

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

(i) $Fe^{3+}_{(aq)}$ ***and*** $I^{-}_{(aq)}$

(ii) $Ag^{+}_{(aq)}$ ***and*** $Cu_{(s)}$

(iii) $Fe^{3+}_{(aq)}$ ***and*** $Cu_{(s)}$

(iv) $Ag_{(s)}$ ***and*** $Fe^{3+}_{(aq)}$

(v) $Br_{2\ (aq)}$ ***and*** $Fe^{2+}_{(aq)}$

Answer:

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

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

Oxidation half reaction: $2\ I^{-}_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^{-}\ ; \qquad E^{\circ} = -0.54V$

Reduction half reaction: $[Fe^{3+}_{(aq)} + e^{-} \rightarrow Fe^{2+}_{(aq)}] \times 2; ; \qquad E^{\circ} = +0.77V$

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

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

(ii) $Ag^{+}_{(aq)}$ and $Cu_{(s)}$

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

Oxidation half reaction: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ e^{-}; \qquad E^{\circ} = -0.34V$

Reduction half reaction: $[Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}] \times 2; \qquad E^{\circ} = +0.80V$

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

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

(iii) $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$

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

Oxidation half reaction: $Cu_{(s)}\ \rightarrow\ Cu^{2+}_{(aq)}\ +\ 2\ e^-;\qquad E^\circ\ =\ -0.34V$

Reduction half reaction: $[Fe^{3+}_{(aq)}\ +\ e^-\ \rightarrow\ Fe^{2+}_{(s)}]\ \times\ 2;\qquad E^\circ\ =\ +0.77V$

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

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

(iv) $Ag_{(s)}$ and $Fe^{3+}_{(aq)}$

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

Oxidation half reaction: $Ag_{(s)}\ \rightarrow\ Ag^+_{(aq)}\ +\ e^-;\qquad E^\circ\ =\ -0.80V$

Reduction half reaction: $Fe^{3+}_{(aq)}\ +\ e^-\ \rightarrow\ Fe^{2+}_{(aq)};\qquad E^\circ\ =\ +0.77V$

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

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

(v) $Br_{2\ (aq)}$ and $Fe^{2+}_{(aq)}$

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

Oxidation half reaction: $[Fe^{2+}_{(aq)}\ \rightarrow\ Fe^{3+}_{(aq)}\ +\ e^-]\ \times\ 2;\qquad E^\circ\ =\ -0.77V$

Reduction half reaction: $Br_{2\ (aq)}\ +\ 2\ e^-\ \rightarrow\ 2\ Br^-_{(aq)};\qquad E^\circ\ =\ +1.09V$

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$$Br_{2\ (s)} \ + \ 2\ Fe^{2+}_{(aq)} \ \rightarrow \ 2\ Br^{-}_{(aq)} \ + \ 2\ Fe^{3+}_{(aq)}; \qquad E^{\circ} \ = \ -0.32V$$

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

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 solution to form Ag^{+} and NO_3^{-} ions.

On electrolysis, either Ag^{+} ion or H_2O molecule can be decreased at cathode. But the reduction

potential of Ag^{+} ions is higher than that of H_2O.

$$Ag^{+}_{(aq)} \ + \ e^{-} \ \rightarrow \ Ag_{(s)}; E^{\circ} \ = \ +0.80V$$

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

Therefore, Ag^{+} ions are decreased at cathode. Same way, Ag metal or H_2O molecules can be

oxidized at anode. But the oxidation potential of Ag is greater than that of H_2O molecules.

$$Ag_{(s)} \ \rightarrow \ Ag^{+}_{(aq)} \ + \ e^{-}; E^{\circ} \ = \ -0.80V$$

$$2\ H_2O_{(l)} \ \rightarrow \ O_{2\ (g)} \ + \ 4\ H^{+}_{(aq)} \ + \ 4\ e^{-}; E^{\circ} \ = \ -1.23V$$

Hence, Ag metal gets oxidized at anode.

(ii) Pt cannot be oxidized very easily. Therefore, at anode, oxidation of water occurs to liberate O_2. At the cathode, Ag^+ ions are decreased and get deposited.

(iii) H_2SO_4 ionizes in aqueous solutions to give H^+ and SO_4^{2-} ions.

$$H_2SO_{4\ (aq)} \rightarrow 2\ H^+_{(aq)} + SO^{2-}_{4(aq)}$$

On electrolysis, either of H_2O molecules or H^+ ions can get decreased at cathode. But the decreased potential of H^+ ions is higher than that of H_2O molecules.

$$2\ H^+_{(aq)} + 2\ e^- \rightarrow H_{2\ (g)}\,;\ E^\circ = 0.0V$$

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

Therefore, at cathode, H^+ ions are decreased to free H_2 gas.

On the other hand, at anode, either of H_2O molecules or SO_4^{2-} ions can be oxidized. But the oxidation of SO_4^{2-} involves breaking of more bonds than that of H_2O molecules. Therefore, SO_4^{2-} ions have lower oxidation potential than H_2O. Hence, H_2O is oxidized at anode to free O_2 molecules.

(iv) In aqueous solutions, $CuCl_2$ ionizes to give Cu^{2+} and Cl^- ions as:

$$CuCl_{2\ (aq)} \rightarrow Cu^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$$

$$CuCl_{2\ (aq)} \rightarrow Cu^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$$

On electrolysis, either of Cu^{2+} ions or H_2O molecules can get decreased at cathode. But the decreased potential of Cu^{2+} is more than that of H_2O molecules.

$$Cu^{2+}_{(aq)} + 2\ e^- \rightarrow Cu_{(aq)}\,;\ E^\circ = +0.34V$$

$$H_2O_{(l)} \;+\; 2\,e^- \;\rightarrow\; H_{2\,(g)} \;+\; 2\,OH^-; E^\circ \;=\; -0.83V$$

Therefore, Cu^{2+} ions are decreased at cathode and get deposited. In the same way, at anode, either of Cl^- or H_2O is oxidized. The oxidation potential of H_2O is higher than that of Cl^-.

$$2\,Cl^-_{(aq)} \;\rightarrow\; Cl_{2\,(g)} \;+\; 2\,e^-;\,E^\circ \;=\; +0.34V$$

$$2\,H_2O_{(l)} \;\rightarrow\; O_{2\,(g)} \;+\; 4\,H^+_{(aq)} \;+\; 4\,e^-;\,E^\circ \;=\; -1.23V$$

But oxidation of H_2O molecules occurs at a lower electrode potential compared to 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.

28. Arrange the given metals in the order in which they displace each other from the solution of their salts.
(i) Al
(ii) Fe
(iii) Cu
(iv) Zn
(v) Mg

Answer:

A metal with stronger reducing power displaces another metal with weaker reducing power from its solution of salt.

The order of the increasing reducing power of the given metals is as given below:

Cu < Fe < Zn < Al < Mg

Therefore, 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 as given below: Mg >Al>Zn> Fe >Cu

29.Given the standard electrode potentials,

K^+ ***/K = −2.93V***

Ag^+ ***/Ag = 0.80V***

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

Mg^{2+} ***/Mg = −2.37V***

Cr^{3+} ***/Cr = −0.74V***

Arrange these metals in their increasing order of reducing power.

Answer:

The reducing agent is stronger as the electrode potential decreases. Hence, the increasing order of the reducing power of the given metals is as given below:

Ag < Hg < Cr < Mg < K

30. Depict the galvanic cell in which the reaction is:

$$Zn_{(s)} \ + \ 2\ Ag^{+}_{(aq)} \ \rightarrow \ Zn^{2+}_{(aq)} \ + \ 2\ Ag_{(s)}$$

Show the following:

(i) Which of the electrode is negatively charged?

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

(iii) Write the individual reaction at each electrode.

Answer:

The galvanic cell corresponding to the given redox reaction can be shown as:

Zn| $Zn^{2+}_{(aq)}$ || $Ag^{+}_{(aq)}$ |Ag

(i) Zn electrode is negatively charged because at this electrode, Zn oxidizes to Zn^{2+} and the leaving electrons accumulate on this electrode.

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

(iii) Reaction at Zn electrode is shown as:

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

Reaction at Ag electrode is shown as:

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