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

(a)

$NaH_2\underline{P}O_4$

(b)

$NaH\underline{S}O_4$

(c)

$H_4\underline{P}_2O_7$

(d)

$K_2\underline{Mn}O_4$

(e)

$Ca\underline{O}_2$

(f)

$Na\underline{B}H_4$

(g)

$H_2\underline{S}_2O_7$

(h)

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

Answer:

(a)

$NaH_2\underline{P}O_4$

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.

(b)

$NaH\underline{S}O_4$

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

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.

(c)

$H_4\underline{P}_2O_7$

$$\overset{+1}{H}_4\overset{x}{P}_2\overset{-2}{O}_7$$

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.

(d)

$K_2\underline{Mn}O_4$

$$\overset{+1}{K}_2\overset{x}{Mn}\overset{-2}{O}_4$$

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.

(e)

$Ca\underline{O}_2$

$$\overset{+2}{Ca}\,\overset{x}{O}_2$$

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.

(f)

$Na\underline{B}H_4$

$$\overset{+1}{Na}\,\overset{x}{B}\overset{-1}{H}_4$$

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.

(g)

$H_2\underline{S}_2O_7$

$$\overset{+1}{H}_2\overset{x}{S}_2\overset{-2}{O}_7$$

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.

(h)

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

$$\overset{+1}{K}\,\overset{3+}{Al}\,(\overset{x}{S}\overset{2-}{O}_4)_2\,.12\overset{+1}{H}_2\overset{-2}{O}$$

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 numbers of the underlined elements in each of the following and how do you rationalise your results?

(a)

(a)

$K\underline{I}_3$

(b)

$H_2\underline{S}_4O_6$

(c)

$\underline{Fe}_3O_4$

(d)

$\underline{C}H_3\underline{C}H_2OH$

(e)

$\underline{C}H_3\underline{C}OOH$

Answer:

(a)

$K\underline{I}_3$

Let x be the oxidation no. of I.

Oxidation no. of K = +1

Then,

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

1 + 3x = 0

x =

−13

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

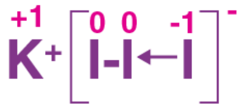
KI_3

.

In

KI_3

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



Therefore, in

$$KI_3$$

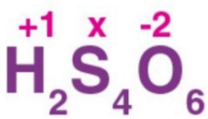
molecule, the oxidation no. of I atoms forming the molecule

$$I_2$$

is 0, while the oxidation no. of I atom, which is forming coordinate bond is -1.

(b)

$$H_2\underline{S}_4O_6$$



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

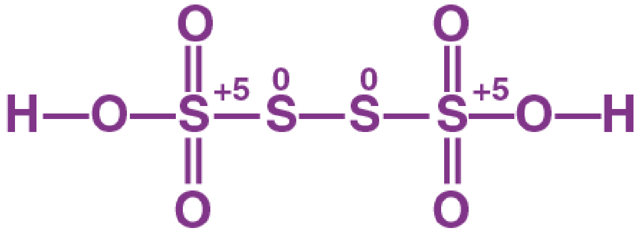
2 + 4x -12 = 0

4x = 10

x =

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

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



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

(c)

$\underline{Fe}_3O_4$

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 the other two atoms of Fe have oxidation no. +3.

$\overset{+2}{Fe}_2O, \overset{-3}{Fe}_2O_3$

(d)

$\underline{C}H_3\underline{C}H_2OH$

$$\overset{x}{C}_2\overset{+1}{H}_6\overset{-2}{O}$$

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.

(e)

$\underline{C}H_3\underline{C}OOH$

$$\overset{x}{C}_2\overset{+1}{H}_4\overset{-2}{O}_2$$

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

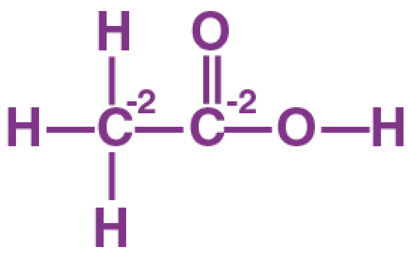
2x + 4 -4 = 0

x = 0

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

CH_3COOH

.



3. Justify that the following reactions are redox reactions:

(a)

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

(b)

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

(c)

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

(d)

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

(e)

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

Answer:

(a)

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

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 a redox reaction.

(b)

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

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 a redox reaction.

(c)

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

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 a redox reaction.

(d)

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

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.

(e)

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

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 a redox reaction.

5. Calculate the oxidation no. 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:

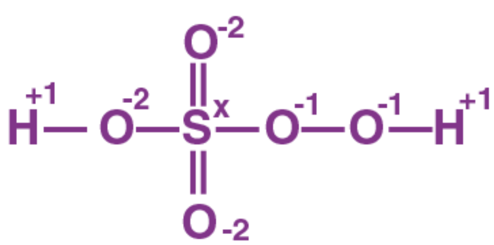
For

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.

For

$Cr_2O_7^{2-}$

Let x be the oxidation no. of Cr.

Oxidation no. of O= -2

Then,

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

2x -14 = -2

x = +6

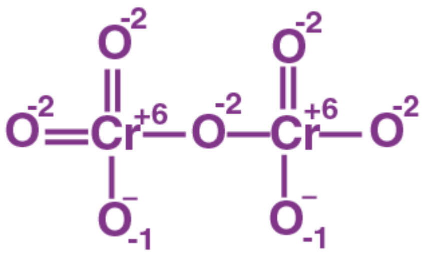
There is no fallacy about the oxidation no. of Cr in

$Cr_2O_7^{2-}$

The structure of

$Cr_2O_7^{2-}$

is as given below.
Each of the two Cr atoms has the oxidation no. of +6.



For

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

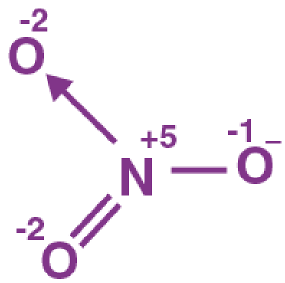
There is no fallacy about the oxidation no. of N in

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:

(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

$HgCl_2$

(b) Nickel (II) sulphate

$NiSO_4$

(c) Tin (IV) oxide

SnO_2

(d) Thallium (I) sulphate

Tl_2SO_4

(e) Iron (III) sulphate

$Fe_2(SO_4)_3$

(f) 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 a 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 oxidation no. +1 and +2.
Therefore,

H_2O_2

can act as a 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:

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

(b)

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

Why it is more appropriate to write these reactions as :

(a)

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

(b)

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

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

Answer:

(a)

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:

[
$$2\ H_2O_{(l)}\ \rightarrow\ 2\ H_{2\ (g)}\ +\ O_{2\ (g)}$$
] × 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 produces water molecules in the photosynthesis process.

The path can be found with the help of radioactive

H_2O^{18}
instead of
H_2O

(b)

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 producing
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 AgF_2 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 a reducing agent and an oxidizing agent, a compound is formed, which has a lower oxidation number if the reducing agent is in excess. A compound is formed which has a 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 an 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?

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

(i) In a neutral medium,

$$OH^-$$

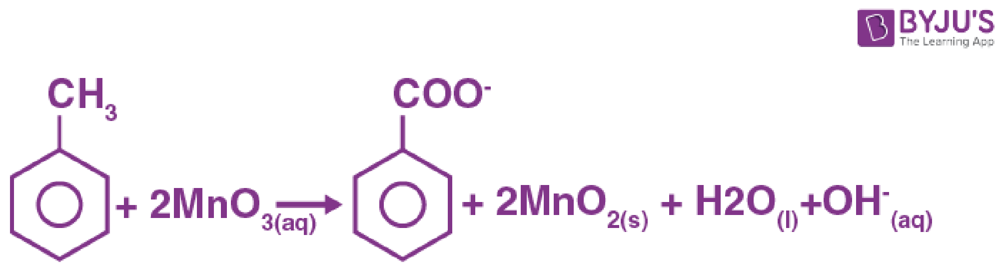
ions are produced in the reaction. Due to that, the cost of adding an acid or a base can be reduced.
(ii)

$$KMnO_4$$

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 a heterogeneous medium. Therefore, in alcohol,

$$KMnO_4$$

and toluene can react at a faster rate.
The redox reaction is as given below:



(b) 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:

(a)

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

(b)

$$HCHO_{(l)}\ +\ 2\ [Ag(NH_3)_2]^+_{(aq)}\ +\ 3\ OH^-_{(aq)}\ \rightarrow\ 2\ Ag_{(s)}\ +\ HCOO^-_{(aq)}\ +\ 4\ NH_{3\ (aq)}\ +\ 2\ H_2O_{(l)}$$

(c)

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

(d)

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

(e)

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

Answer:

(a)

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

$C_6H_6O_2$
=> Oxidized substance

AgBr => Reduced substance

AgBr =>Oxidizing agent

$C_6H_6O_2$
=> Reducing agent

(b)

$$HCHO_{(l)} + 2\,[Ag(NH_3)_2]^{+}_{(aq)} + 3\,OH^{-}_{(aq)} \rightarrow 2\,Ag_{(s)} + HCOO^{-}_{(aq)} + 4\,NH_{3\,(aq)} + 2\,H_2O_{(l)}$$

HCHO => Oxidized substance

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

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

HCHO=> Reducing agent

(c)

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

HCHO => Oxidized substance

Cu^{2+}
=> Reduced substance

Cu^{2+}
=> Oxidizing agent

HCHO => Reducing agent

(d)

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

N_2H_4
=> Oxidized substance

H_2O_2
=> Reduced substance

H_2O_2
=> Oxidizing agent

N_2H_4
=> Reducing agent

(e)

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

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, 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 reacts 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 about the compound

Na_4XeO_6
(of which
XeO^{4-}_{6}
is a part) can be drawn 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^{4-}_{6}
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.

(a)

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

(b)

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

(c)

$$C_6H_5CHO_{(l)} + 2\,[Ag(NH_3)_2]^+_{(aq)} + 3\,OH^-_{(aq)} \rightarrow C_6H_5COO^-_{(aq)} + 2\,Ag_{(s)} + 4\,NH_{3\,(aq)} + 2\,H_2O_{(l)}$$

(d)

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

What inference do you draw about the behavior of

Ag^+
and
Cu^{2+}
from these reactions?

Answer:

Ag^{+}
and
Cu^{2+}
behave as oxidizing agents 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+}
.

(a)

$MnO_{4\ (aq)}^{-} + I_{(aq)}^{-} \rightarrow MnO_{2\ (s)} + I_{2\ (s)}$
(Basic medium)

(b)

$MnO_{4\ (aq)}^{-} + SO_{2\ (g)} \rightarrow Mn_{(aq)}^{2+} + H_2SO_4^{-}$
(Acidic medium)

(c)

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

(d)

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

Answer:

(a)

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

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 the 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 sides by multiplying the oxidation half-reaction by 3 and the 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:

$$6\,I_{(aq)}^{-} + 2\,MnO_{4\ (aq)}^{-} + 4\,H_2O_{(l)} \rightarrow 3\,I_{2\ (s)} + 2\,MnO_{2\ (s)} + 8\,OH_{(aq)}^{-}$$

(b)

$$MnO_{4\ (aq)}^{-} + SO_{2\ (g)} \rightarrow Mn_{(aq)}^{2+} + H_2SO_4^{-}$$

Step 1

Similar to (i), the 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:

$$2\,MnO_4^-{}_{(aq)} + 5\,SO_{2\,(g)} + 2\,H_2O_{(l)} + H^+_{(aq)} \rightarrow 2\,Mn^{2+}_{(aq)} + 5\,HSO_4^-{}_{(aq)}$$

(c)

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

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:

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

(d)

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

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:

$$Cr^{2-}_2O_{7\ (aq)} + 3\ SO_{2\ (g)} + 2\ H^+_{(aq)} \rightarrow 2\ Cr^{3+}_{(aq)} + 3\ SO^{2-}_{4\ (aq)} + H_2O_{(l)}$$

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.

(a)

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

(b)

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

(c)

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

Answer:

(a) 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 a reducing agent as well as an oxidizing 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:

$$P_{4\,(s)} + 12\,OH_{(aq)}^- \rightarrow 4\,HPO_{2\,(aq)}^- + 4\,H_2O_{(l)} + 8\,e^-$$
–––-(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:

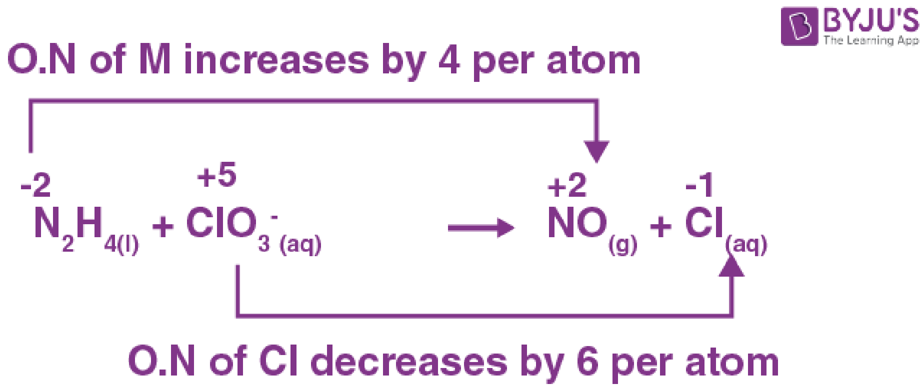
$$P_{4\ (s)}\ +\ 12\ H_2O_{(l)}\ +\ 12\ e^{-}\ \rightarrow\ 4\ PH_{3\ (g)}\ +\ 12\ OH^{-}_{(aq)}$$

––-- (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:

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

(b)



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:

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

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

Now, multiply 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:

$$3\ N_2H_{4\ (l)}\ +\ 4\ ClO_{3\ (aq)}^{-}\ \rightarrow\ 6\ NO_{(g)}\ +\ 4\ Cl_{(aq)}^{-}$$

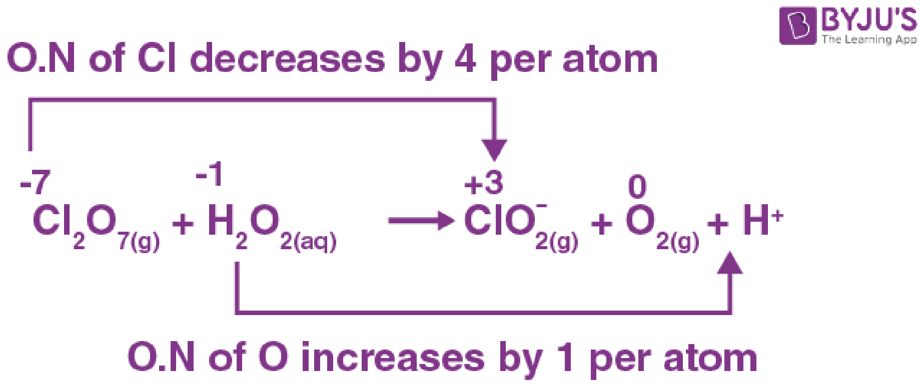
– Add 6

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.

(c)



The Oxidation no. of Cldecreases from +7 in

Cl_2O_7
to +3 in
ClO_2^-
.

The oxidation no. of increases 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:

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

$Cl_2O_{7\,(g)} \rightarrow 2\,ClO_{2\,(aq)}^-$

– 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 the alkaline decomposition of cyanogens is a disproportionation reaction.

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 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:

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

$----$ (1)

The reduction half-reaction:

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

Add 1 electron to balance oxidation no. :

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

Add equations (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, Ne, I and F

(a) Identify the element that exhibits only negative oxidation.

(b) Identify the element that exhibits only positive oxidation.

(c) Identify the element that exhibits both negative and positive oxidation states.

(d) Identify the element that exhibits neither negative nor positive oxidation state?

Answer:

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

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

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

(d) 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 equations (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:

(a) Select the possible non–metals that can show disproportionation reaction?

(b) Select three metals that show 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
$\frac{160\ \times\ 10}{68}$
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
$\frac{120 \times 20}{160}$
g of N = 15 g of NO.

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

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

$Fe^{3+}_{(aq)}$
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)

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

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

Oxidation half reaction:

$2\ I^{-}_{(aq)}\ \rightarrow\ I_{2\ (s)}\ +\ 2\ e^{-}$
;
$E^{o}\ =\ -0.54V$

Reduction half reaction:

$[Fe^{3+}_{(aq)}\ +\ e^{-}\ \rightarrow\ Fe^{2+}_{(aq)}]\ \times\ 2;$
;
$E^{o}\ =\ +0.77V$

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

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

(b)

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

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

Oxidation half reaction:

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

Reduction half reaction:

$$[Ag^{+}_{(aq)}\ +\ e^{-}\ \rightarrow\ Ag_{(s)}]\ \times\ 2$$
;
$$E^{o}\ =\ +0.80V$$

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

E^{o}

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

(c)

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

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

Oxidation half reaction:

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

Reduction half reaction:

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

;

$E^{o} = +0.77V$

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

;

$E^{o} = +0.43V$

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

(d)

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

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

Oxidation half reaction:

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

;

$E^{o} = -0.80V$

Reduction half reaction:

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

;

$E^{o} = +0.77V$

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

;

$E^{o} = -0.03V$

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

(e)

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

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

Oxidation half reaction:

$[Fe^{2+}_{(aq)}\ \rightarrow\ Fe^{3+}_{(aq)}\ +\ e^{-}]\ \times\ 2$
;
$E^{o}\ =\ -0.77V$

Reduction half reaction:

$Br_{2\ (aq)}\ +\ 2\ e^{-}\ \rightarrow\ 2\ Br^{-}_{(aq)}$
;
$E^{o}\ =\ +1.09V$

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

E^{o}
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
.

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

,

$$E^o = +0.80V$$

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

;

$$E^o = -0.83V$$

Therefore,

Ag^+
ions are decreased at the cathode. Same way, Ag metal or
H_2O
molecules can be oxidized at the anode. But the oxidation potential of Ag is greater than that of
H_2O
molecules.

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

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

Hence, Ag metal gets oxidized at the anode.

(ii) Pt cannot be oxidized very easily. Therefore, at the 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.

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

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

Therefore, at cathode,

H^+
ions are decreased to free
H_2
gas.

On the other hand, at the 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.

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

;
$E^o = +0.34V$

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

Therefore,

Cu^{2+}
ions are decreased at the cathode and get deposited. In the same way, at the anode, either of
Cl^-

or
H_2O
is oxidized. The oxidation potential of
H_2O
is higher than that of
Cl^-

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

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

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.

Al, Fe, Cu, Zn, 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:

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

Further show:

(i) which of the electrode is negatively charged?

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

(iii) 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:

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