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

(a) \( \text{NaH}_2\text{PO}_4 \)

(b) \( \text{NaHSO}_4 \)

(c) \( \text{H}_4\text{P}_2\text{O}_7 \)

(d) \( \text{K}_2\text{Mn}_2\text{O}_4 \)

(e) \( \text{CaO}_2 \)

(f) \( \text{NaBH}_4 \)

(g) \( \text{H}_2\text{S}_2\text{O}_7 \)

(h) \( \text{KAl(SO}_4\text{)}_2\cdot12\text{H}_2\text{O} \)

Answer:

(a) \( \text{NaH}_2\text{PO}_4 \)

Let \( x \) be the oxidation no. of P.

Oxidation no. of Na = +1

Oxidation no. of H = +1

Oxidation no. of O = -2
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) \(NaHSO_4\)

\[Na^{+1} H^{+1} S^{x} O^{-2}_4\]

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) \(H_4P_2O_7\)

\[H^{+1}_4 P^{x}_2 O^{-2}_7\]

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) \( K_2MnO_4 \)

\[
\begin{array}{c}
+1 \\
K_2 \\
+2 \\
MnO_4
\end{array}
\]

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) \( CaO_2 \)

\[
\begin{array}{c}
+2 \\
Ca \\
+x \\
O_2
\end{array}
\]

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) \( \text{NaBH}_4 \)

\[
\begin{array}{c}
\text{Na} & \text{B} & \text{H}_4 \\
+1 & x & -1 \\
\end{array}
\]

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) \( \text{H}_2\text{S}_2\text{O}_7 \)

\[
\begin{array}{c}
\text{H} & \text{S}_2 & \text{O}_7 \\
+1 & x & -2 \\
\end{array}
\]

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) \( \text{KAl(SO}_4\text{)}_2.12\text{H}_2\text{O} \)
2. What are the oxidation number of the underlined elements in each of the following and how do you rationalise your results?

(a) \( K \text{I}_3 \)

(b) \( H_2S_4O_6 \)

(c) \( Fe_3O_4 \)
(d) \( \text{CH}_3\text{CH}_2\text{OH} \)

(e) \( \text{CH}_3\text{C} = \text{O} \text{OH} \)

**Answer:**

(a) \( KI_3 \)

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.

![Image of KI3 structure]

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) \( H_2S_4O_6 \)

Let \( x \) be the oxidation no. of S.

Oxidation no. of H = +1

Oxidation no. of O = -2
Then,
\[2(+1) + 4(x) + 6(-2) = 0\]
\[2 + 4x -12 = 0\]
\[4x = 10\]
\[x = \frac{5}{2}\]

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

\[\text{H—O—S}^{+5} \text{O}^0 \text{S—S—S}^{+5} \text{—O—H}\]

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

(c) \(Fe_3O_4\)

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 other two atoms of Fe has oxidation no. +3.

\[Fe_2O, Fe_2O_3\]
(d) \( \text{CH}_3\text{CH}_2\text{OH} \)

\[
\begin{array}{c}
\text{C}_2\text{H}_6\text{O} \\
\begin{array}{c}
x + 1 \\
-2
\end{array}
\end{array}
\]

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) \( \text{CH}_3\text{COOH} \)

\[
\begin{array}{c}
\text{C}_2\text{H}_4\text{O}_2 \\
\begin{array}{c}
x + 1 \\
-2
\end{array}
\end{array}
\]

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 \( \text{CH}_3\text{COOH} \).
3. Justify that the following reactions are redox reactions:

(a) \( \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(g) \)

(b) \( \text{Fe}_2\text{O}_3(s) + 3 \text{CO}(g) \rightarrow 2 \text{Fe}(s) + 3 \text{CO}_2(g) \)

(c) \( 4 \text{BCl}_3(g) + 3 \text{LiAlH}_4(s) \rightarrow 2 \text{B}_2\text{H}_6(g) + 3 \text{LiCl}(s) + 3 \text{AlCl}_3(s) \)

(d) \( 2 \text{K}(s) + \text{F}_2(g) \rightarrow 2 \text{K} + \text{F}(s) \)

(e) \( 4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(g) \)

Answer:

(a) \( \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(g) \)

Oxidation no. of Cu and O in \( \text{CuO} \) is +2 and -2 respectively.

Oxidation no. of \( \text{H}_2 \) is 0.

Oxidation no. of Cu is 0.

Oxidation no. of H and O in \( \text{H}_2\text{O} \) is +1 and -2 respectively.

The oxidation no. of Cu decreased from +2 in \( \text{CuO} \) to 0 in Cu. That is \( \text{CuO} \) is reduced to Cu.

The oxidation no. of H increased from 0 to +1 in \( \text{H}_2 \). That is \( \text{H}_2 \) is oxidized to \( \text{H}_2\text{O} \).
Therefore, the reaction is redox reaction.

(b) \( Fe_2O_3 \ (s) + 3 \ \text{CO} \ (g) \rightarrow 2 \ \text{Fe} \ (s) + 3 \ \text{CO}_2 \ (g) \)

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.

(c) \( 4 \ \text{BCl}_3 \ (g) + 3 \ \text{LiAlH}_4 \ (s) \rightarrow 2 \ \text{B}_2\text{H}_6 \ (g) + 3 \ \text{LiCl} \ (s) + 3 \ \text{AlCl}_3 \ (s) \)

In the above reaction,

Oxidation no. of B and Cl in \( \text{BCl}_3 \) is +3 and -1 respectively.

Oxidation no. of Li, Al and H in \( \text{LiAlH}_4 \) is +1, +3 and -1 respectively.

Oxidation no. of B and H in \( \text{B}_2\text{H}_6 \) is -3 and +1 respectively.

Oxidation no. of Li and Cl in \( \text{LiCl} \) is +1 and -1 respectively.

Oxidation no. of Al and Cl in \( \text{AlCl}_3 \) 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.

\[(d) \quad 2 \text{K}_\text{(s)} + F_2(\text{g}) \rightarrow 2 \text{K} + F_\text{(s)}\]

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.

\[(e) \quad 4 \text{NH}_3(\text{g}) + 5 O_2(\text{g}) \rightarrow 4 \text{NO}_2(\text{g}) + 6 H_2O(\text{g})\]

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 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,
\[2(x) + 7(-2) = -2\]
\[2x - 14 = -2\]
\[2x = +6\]
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.
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

\[ \text{HgCl}_2 \]

(b) Nickel (II) sulphate

\[ \text{NiSO}_4 \]

(c) Tin (IV) oxide

\[ \text{SnO}_2 \]

(d) Thallium (I) sulphate

\[ \text{Tl}_2\text{SO}_4 \]

(e) Iron (III) sulphate

\[ \text{Fe}_2(\text{SO}_4)_3 \]

(f) Chromium (III) oxide

\[ \text{Cr}_2\text{O}_3 \]
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:

<table>
<thead>
<tr>
<th>Compounds</th>
<th>Oxidation no. of carbon</th>
</tr>
</thead>
<tbody>
<tr>
<td>$CH_2Cl_2$</td>
<td>0</td>
</tr>
<tr>
<td>$HC\equiv CH$</td>
<td>-1</td>
</tr>
<tr>
<td>$ClC\equiv CCl$</td>
<td>+1</td>
</tr>
<tr>
<td>$CH_3Cl$</td>
<td>-2</td>
</tr>
<tr>
<td>$CHCl_3, CO$</td>
<td>+2</td>
</tr>
<tr>
<td>$H_3C – CH_3$</td>
<td>-3</td>
</tr>
<tr>
<td>$Cl_3C – CCl_3$</td>
<td>+3</td>
</tr>
<tr>
<td>$CH_4$</td>
<td>-4</td>
</tr>
<tr>
<td>$CCl_4, CO_2$</td>
<td>+4</td>
</tr>
</tbody>
</table>
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 (\(S\text{O}_2\)) 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 the 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:

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

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

Why it is more appropriate to write these reactions as:

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

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

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

Step 1:

\[ 2 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_2(g) + \text{O}_2(g) \]

Step 2:

The \( \text{H}_2 \) produced in earlier step reduces \( \text{CO}_2 \), thus produce glucose and water.

\[ 6 \text{CO}_2(g) + 12 \text{H}_2(g) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{H}_2\text{O}(l) \]

The net reaction is as given below:

\[ [2 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_2(g) + \text{O}_2(g)] \times 5 \]

\[ 6 \text{CO}_2(g) + 12 \text{H}_2(g) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{H}_2\text{O}(l) \]

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 \( \text{H}_2\text{O}^{18} \) instead of \( \text{H}_2\text{O} \).

(b)

Step 1:

\( \text{O}_2 \) is produced from each of the reactants \( \text{O}_3 \) and \( \text{H}_2\text{O}_2 \). That is the reason \( \text{O}_2 \) is written two times.
10. The compound AgF$_2$ is an unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why?

Answer:

The oxidation no. of Ag in $\text{AgF}_2$ is +2. But, +2 is a very unstable oxidation no. of Ag. Hence, when $\text{AgF}_2$ is formed, silver accepts an electron and forms $\text{Ag}^+$. This decreases the oxidation no. of Ag from +2 to +1. +1 state is more stable. Therefore, $\text{AgF}_2$ 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) \( \text{P}_4 \) and \( \text{F}_2 \) are reducing and oxidizing agent respectively.

In an excess amount of \( \text{P}_4 \) is reacted with \( \text{F}_2 \), then \( \text{PF}_3 \) would be produced, where the oxidation no. of \( \text{P} \) is +3.

\[
\text{P}_4 \text{(excess)} + \text{F}_2 \rightarrow \text{PF}_3
\]

If \( \text{P}_4 \) is reacted with excess of \( \text{F}_2 \), then \( \text{PF}_5 \) would be produced, where the oxidation no. of \( \text{P} \) is +5.

\[
\text{P}_4 + \text{F}_2 \text{(excess)} \rightarrow \text{PF}_5
\]

(ii) \( \text{K} \) and \( \text{O}_2 \) acts as a reducing agent and oxidizing agent respectively.

If an excess of \( \text{K} \) reacts with \( \text{O}_2 \), it produces \( \text{K}_2\text{O} \). Here, the oxidation number of \( \text{O} \) is -2.

\[
4 \text{K} \text{(excess)} + \text{O}_2 \rightarrow 2 \text{K}_2\text{O}^{-2}
\]

If \( \text{K} \) reacts with an excess of \( \text{O}_2 \), it produces \( \text{K}_2\text{O}_2 \), where the oxidation number of \( \text{O} \) is -1.

\[
2 \text{K} + \text{O}_2 \text{(excess)} \rightarrow \text{K}_2\text{O}_2^{-1}
\]

(iii) \( \text{C} \) and \( \text{O}_2 \) 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?

(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 heterogeneous medium. Therefore, in alcohol, $KMnO_4$ and toluene can react at a faster rate.

The redox reaction is as given below:
13. Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:

(a) \[ 2 \text{AgBr}_\text{(s)} + \text{C}_6\text{H}_6\text{O}_2\text{(aq)} \rightarrow 2 \text{Ag}_\text{(s)} + 2 \text{HBr}_\text{(aq)} + \text{C}_6\text{H}_4\text{O}_2\text{(aq)} \]

(b) \[ \text{HCHO}(l) + 2 [\text{Ag(NH}_3)_2]^+ + 3 \text{OH}^- \rightarrow 2 \text{Ag}_\text{(s)} + \text{HCOO}^- + 4 \text{NH}_3 + 2 \text{H}_2\text{O} \]

(c) \[ \text{HCHO}(l) + 2 \text{Cu}^{2+} + 5 \text{OH}^- \rightarrow \text{Cu}_2\text{O}_\text{(s)} + \text{HCOO}^- + 3 \text{H}_2\text{O} \]
(d) \( \text{N}_2\text{H}_4(l) + 2 \text{H}_2\text{O}_2(l) \rightarrow \text{N}_2(g) + 4 \text{H}_2\text{O}(l) \)

(e) \( \text{Pb}(s) + \text{PbO}_2(s) + 2 \text{H}_2\text{SO}_4(aq) \rightarrow 2 \text{PbSO}_4(aq) + 2 \text{H}_2\text{O}(l) \)

Answer:

(a) \( 2 \text{AgBr}(s) + \text{C}_6\text{H}_6\text{O}_2(aq) \rightarrow 2 \text{Ag}(s) + 2 \text{HBr}(aq) + \text{C}_6\text{H}_4\text{O}_2(aq) \)

\( \text{C}_6\text{H}_6\text{O}_2 \) => Oxidized substance

\( \text{AgBr} \) => Reduced substance

\( \text{AgBr} \) => Oxidizing agent

\( \text{C}_6\text{H}_6\text{O}_2 \) => Reducing agent

(b) \( \text{HCHO}(l) + 2 [\text{Ag(NH}_3)_2]^+(aq) + 3 \text{OH}^-(aq) \rightarrow 2 \text{Ag}(s) + \text{HCOO}^-(aq) + 4 \text{NH}_3(aq) + 2 \text{H}_2\text{O}(l) \)

\( \text{HCHO} \) => Oxidized substance

\( [\text{Ag(NH}_3)_2]^+ \) => Reduced substance

\( [\text{Ag(NH}_3)_2]^+ \) => Oxidizing agent

\( \text{HCHO} \) => Reducing agent

(c) \( \text{HCHO}(l) + 2 \text{Cu}^{2+}(aq) + 5 \text{OH}^-(aq) \rightarrow \text{Cu}_2\text{O}(s) + \text{HCOO}^-(aq) + 3 \text{H}_2\text{O}(l) \)

\( \text{HCHO} \) => Oxidized substance

\( \text{Cu}^{2+} \) => Reduced substance

\( \text{Cu}^{2+} \) => Oxidizing agent
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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:**

\[2 S_2O_3^{2-} (aq) + I_2 (s) \rightarrow S_4O_6^{2-} (aq) + 2 I^- (aq)\]

\[S_2O_3^{2-} (aq) + 2 Br_2 (l) + 5 H_2O (l) \rightarrow 2 SO_4^{2-} (aq) + 4 Br^- (aq) + 10 H^+ (aq)\]

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

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

**Answer:**

\[ F_2 \text{ can oxidize } Cl^- \text{ to } Cl_2, \text{ Br}^- \text{ to } Br_2, \text{ and } I^- \text{ to } I_2 \text{ as:} \]

\[ F_2(aq) + 2 Cl^-(s) \rightarrow 2 F^{(aq)} + Cl_2(g) \]

\[ F_2(aq) + 2 Br^-(aq) \rightarrow 2 F^{(aq)} + Br_2(l) \]
\[ \text{F}_2 \text{(aq)} + 2 \text{I}_2 \text{(aq)} \rightarrow 2 \text{F}^- \text{(aq)} + \text{I}_2 \text{(s)} \]

But, \( \text{Cl}_2 \), \( \text{Br}_2 \), and \( \text{I}_2 \) cannot oxidize \( \text{F}^- \) to \( \text{F}_2 \). The oxidizing power of halogens increases in the order as given below:

\( \text{I}_2 < \text{Br}_2 < \text{Cl}_2 < \text{F}_2 \)

Therefore, fluorine is the best oxidant among halogens.

\( \text{HI} \) and \( \text{HBr} \) can reduce \( \text{H}_2\text{SO}_4 \) to \( \text{SO}_2 \), but \( \text{HCl} \) and \( \text{HF} \) cannot. Hence, \( \text{HI} \) and \( \text{HBr} \) are stronger reductants compared to \( \text{HCl} \) and \( \text{HF} \).

\[ 2 \text{HI} + \text{H}_2\text{SO}_4 \rightarrow \text{I}_2 + \text{SO}_2 + 2 \text{H}_2\text{O} \]

\[ 2 \text{HBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Br}_2 + \text{SO}_2 + 2 \text{H}_2\text{O} \]

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

\[ 4 \text{I}^- \text{(aq)} + 2 \text{Cu}^{2+} \text{(aq)} \rightarrow \text{Cu}_2\text{I}_2 \text{(s)} + \text{I}_2 \text{(aq)} \]

Therefore, hydrochloric acid is the best reductant among hydrohalic compounds.

Hence, the reducing power of hydrohalic acids increases as given below:

\[ \text{HF} < \text{HCl} < \text{HBr} < \text{HI} \]

16. Why does the following reaction occur?

\[ \text{XeO}_4^{2-} \text{(aq)} + 2 \text{F}^- \text{(aq)} + 6 \text{H}^+ \text{(aq)} \rightarrow \text{XeO}_3 \text{(g)} + \text{F}_2 \text{(g)} + 3 \text{H}_2\text{O} \text{(l)} \]
What conclusion about the compound \( \text{Na}_4 \text{XeO}_6 \) (of which \( \text{XeO}_6^{4-} \) is a part) can be drawn from the reaction?

Answer:

\[
\text{XeO}_6^{4-} + 2 \text{F}^- + 6 \text{H}^+ \rightarrow \text{XeO}_3 (g) + \text{F}_2 (g) + 3 \text{H}_2 \text{O} (l)
\]

The oxidation no. of Xe reduces from +8 in \( \text{XeO}_6^{4-} \) to +6 in \( \text{XeO}_3 \).

The oxidation no. of F increases from -1 in \( \text{F}^- \) to 0 in \( \text{F}_2 \).

Hence, \( \text{XeO}_6^{4-} \) is reduced on the other hand \( \text{F}^- \) is oxidized. As \( \text{Na}_2 \text{XeO}_6^{4-} \) (or \( \text{XeO}_6^{4-} \)) is a stronger oxidizing agent compared to \( \text{F}_2 \), this reaction occurs.

17. Consider the reactions:

(a) \( \text{H}_3\text{PO}_2 (aq) + 4 \text{AgNO}_3 (aq) + 2 \text{H}_2\text{O} (l) \rightarrow \text{H}_3\text{PO}_4 (aq) + 4 \text{Ag(s)} + 4 \text{HNO}_3 (aq) \)

(b) \( \text{H}_3\text{PO}_2 (aq) + 2 \text{CuSO}_4 (aq) + 2 \text{H}_2\text{O} (l) \rightarrow \text{H}_3\text{PO}_4 (aq) + 2 \text{Cu(s)} + \text{H}_2\text{SO}_4 (aq) \)

(c) \( \text{C}_6\text{H}_5\text{CHO} (l) + 2 [\text{Ag(NH}_3)_2]^+ + 3 \text{OH}^- (aq) \rightarrow \text{C}_6\text{H}_5\text{COO}^- (aq) + 2 \text{Ag(s)} + 4 \text{NH}_3 (aq) + 2 \text{H}_2\text{O} (l) \)
(d) \( C_6H_5CHO \text{(l)} + 2 Cu^{2+} (aq) + 5 OH^{-} (aq) \rightarrow \text{No change is observed} \)

What inference do you draw about the behavior of \( Ag^{+} \) and \( Cu^{2+} \) from these reactions?

Answer:

\( Ag^{+} \) and \( Cu^{2+} \) 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:

(a) \( MnO_4^{-} (aq) + I^{-} (aq) \rightarrow MnO_2 (s) + I_{2} (s) \) (Basic medium)

(b) \( MnO_4^{-} (aq) + SO_2 (g) \rightarrow Mn^{2+} (aq) + H_2SO_4 \) (Acidic medium)

(c) \( H_2O_2 (aq) + Fe^{2+} (aq) \rightarrow Fe^{3+} (aq) + H_2O (l) \) (Acidic medium)

(d) \( Cr_2O_7^{2-} (aq) + SO_2 (g) \rightarrow Cr^{3+} (aq) + SO_2^{-} (aq) \) (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 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.

\[
MnO_4^-_{(aq)} + 2 H_2O + 3 e^- \rightarrow MnO_2_{(aq)} + 4 OH^-
\]

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 5
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) $$MnO_4^-_{(aq)} + SO_2 (g) \rightarrow Mn^{2+}_{(aq)} + H_2SO_4^-$$

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:

$$2 \, MnO_4^-_{(aq)} + 5 \, SO_2 (g) + 2 \, H_2O_{(l)} + 5 \, H^+_{(aq)} \rightarrow 2 \, Mn^{2+}_{(aq)} + 5 \, HSO_4^-_{(aq)}$$

(c) $$H_2O_2_{(aq)} + Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + H_2O_{(l)}$$

Step 1
Similar to (i), oxidation half reaction is:
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) \( \text{P}_4 \ (s) \ + \ \text{OH}_- \ (aq) \ \rightarrow \ \text{PH}_3 \ (g) \ + \ \text{HPO}_2^- \ (aq) \)
(b) \( \text{N}_2\text{H}_4(l) + \text{ClO}_3\text{aq} \rightarrow \text{NO}_g + \text{Cl}_g \)

(c) \( \text{Cl}_2\text{O}_7(g) + \text{H}_2\text{O}_2\text{aq} \rightarrow \text{ClO}_2\text{aq} + \text{O}_2(g) + \text{H}_+\text{aq} \)

Answer:

(a) The Oxidation no. of P reduces from 0 in \( \text{P}_4 \) to \(-3\) in \( \text{PH}_3 \)

The oxidation no. of P increases from 0 in \( \text{P}_4 \) to \(+2\) in \( \text{HPO}_2^- \). Therefore, \( \text{P}_4 \) behaves both as a reducing agent as well as an oxidizing agent in the reaction.

Ion – electron method:

– The oxidation half reaction:

\[ \text{P}_4(s) \rightarrow 4 \text{HPO}_2^-\text{aq} \]

– Balance atom P:

\[ \text{P}_4(s) \rightarrow 4 \text{HPO}_2^-\text{aq} \]

– Add 8 electrons to balance oxidation no.

\[ \text{P}_4(s) \rightarrow 4 \text{HPO}_2^-\text{aq} + 8\text{e}^- \]

– Add \( 12 \text{OH}^- \) to balance the charge:

\[ \text{P}_4(s) + 12 \text{OH}^-\text{aq} \rightarrow 4 \text{HPO}_2^-\text{aq} + 8\text{e}^- \]

– Add 4 \( \text{H}_2\text{O} \) to balance H and O atoms:

\[ \text{P}_4(s) + 12 \text{OH}^-\text{aq} \rightarrow 4 \text{HPO}_2^-\text{aq} + 4 \text{H}_2\text{O}_1 + 8\text{e}^- \]

– The reduction half reaction:
\[ P_4(s) \rightarrow PH_3(g) \]

- Balance atom P:

\[ P_4^0(s) \rightarrow 4P^{3-}H_3(g) \]

- Add 12 electrons to balance oxidation no.

\[ P_4(s) + 12e^- \rightarrow 4PH_3(g) \]

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

\[ P_4(s) + 12e^- \rightarrow 4PH_3(g) + 12OH^-(aq) \]

- Add 12 $H_2O$ to balance H and O atoms:

\[ P_4(s) + 12H_2O(l) + 12e^- \rightarrow 4PH_3(g) + 12OH^-(aq) \quad \text{(2)} \]

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:

\[ 5P_4(s) + 12H_2O(l) + 12HO^-(aq) \rightarrow 8PH_3(g) + 12HPO_2^-(aq) \]

(b)

**O.N. of M increases by 4 per atom**

\[ \text{N}_2\text{H}_4(l) + \text{ClO}_3^-(aq) \rightarrow \text{NO}(g) + \text{Cl}^-(aq) \]

**O.N. of Cl decreases by 6 per atom**
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:

  $$N_2H_4 (l) + 8 OH^-_{(aq)} \rightarrow 2 NO(g) + 6 H_2O(l) + 8 e^- \quad \text{--- (1)}$$

- The reduction half reaction:

  $$ClO_3^-_{(aq)} \rightarrow Cl^-_{(aq)}$$
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- 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:

\[ ClO_3^{-}_{(aq)} + 3 H_2O_{(l)} + 6 e^- \rightarrow Cl^{-}_{(aq)} + 6 OH^-_{(aq)} \]

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 \times 4 = 8
- Increment in the oxidation no. of Cl = 1 \times 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 \( N_2H_4(l) + 4 ClO_3^{-}(aq) \rightarrow 6 NO(g) + 4 Cl^{-}(aq) \)

- 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 Cl decreases by 4 per atom from +7 in \( ClO_7 \) to +3 in \( ClO_2^- \).

The oxidation no. of O increases from -1 in \( H_2O_2 \) to 0 in \( O_2 \).

Therefore, \( H_2O_2 \) behaves as a reducing agent while \( ClO_7 \) 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:

$$H_2O_2 \ (aq) + 2 \ OH^- \ (aq) \rightarrow O_2 \ (g) + 2 \ H_2O \ (l) + 2 \ e^- \quad \text{--- (1)}$$

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:
\[ \text{Cl}_2\text{O}_7 \quad \text{(g)} \quad + \quad 4 \ H_2\text{O}_2 \quad \text{(aq)} \quad + \quad 2 \ \text{OH}^- \quad \text{(aq)} \quad \rightarrow \quad 2 \ \text{ClO}_2^- \quad \text{(aq)} \quad + \quad 4 \ \text{O}_2 \quad \text{(g)} \quad + \quad 5 \ \text{H}_2\text{O} \quad \text{(l)} \]

Oxidation number method:

- Reduction in the oxidation no. of \( \text{Cl}_2\text{O}_7 = 4 \times 2 = 8 \)

- Increment in the oxidation no. of \( \text{H}_2\text{O}_2 = 2 \times 1 = 2 \)

Multiply \( \text{H}_2\text{O}_2 \) by 4 and \( \text{O}_2 \) by 4 to balance the reduction and increment of the oxidation no.: 

\[ 3 \ \text{N}_2\text{H}_4 \quad \text{(l)} \quad + \quad 4 \ \text{ClO}_3^- \quad \text{(aq)} \quad \rightarrow \quad \text{NO} \quad \text{(g)} \quad + \quad \text{Cl}^- \quad \text{(aq)} \]

- Balance Cl and n atoms:

\[ \text{Cl}_2\text{O}_7 \quad \text{(g)} \quad + \quad 4 \ \text{H}_2\text{O}_2 \quad \text{(aq)} \quad \rightarrow \quad 2 \ \text{ClO}_2^- \quad \text{(aq)} \quad + \quad 4 \ \text{O}_2 \quad \text{(g)} \]

- Add 3 \( \text{H}_2\text{O} \) to balance 0 atoms:

\[ \text{Cl}_2\text{O}_7 \quad \text{(g)} \quad + \quad 4 \ \text{H}_2\text{O}_2 \quad \text{(aq)} \quad \rightarrow \quad 2 \ \text{ClO}_2^- \quad \text{(aq)} \quad + \quad 4 \ \text{O}_2 \quad \text{(g)} \quad + \quad 3 \ \text{H}_2\text{O} \quad \text{(l)} \]

- Add 2 \( \text{OH}^- \) and 2 \( \text{H}_2\text{O} \) to balance H atoms:

\[ \text{Cl}_2\text{O}_7 \quad \text{(g)} \quad + \quad 4 \ \text{H}_2\text{O}_2 \quad \text{(aq)} \quad 2 \ \text{OH}^- \quad \text{(aq)} \quad \rightarrow \quad 2 \ \text{ClO}_2^- \quad \text{(aq)} \quad + \quad 4 \ \text{O}_2 \quad \text{(g)} \quad + \quad 5 \ \text{H}_2\text{O} \quad \text{(l)} \]

This is the required reaction equation.

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

\[ (\text{CN})_2 \quad \text{(g)} \quad + \quad 2 \ \text{OH}^- \quad \text{(aq)} \quad \rightarrow \quad \text{CN}^- \quad \text{(aq)} \quad + \quad \text{CNO}^- \quad \text{(aq)} \quad + \quad \text{H}_2\text{O} \quad \text{(l)} \]
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, \(y = 2\)

\(CNO^-\)

\(y - 3 \cdot 2 = -1\)
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. 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:

\[ Mn^{3+} \rightarrow Mn^{2+} + MnO_2(s) + H^+ \]

The oxidation half reaction:

\[ Mn^{3+} \rightarrow MnO_2(s) \]

Add 1 electron to balance the oxidation no.:

\[ Mn^{3+} \rightarrow MnO_2(s) + e^- \]

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

\[ Mn^{3+} \rightarrow MnO_2(s) + e^- + 4 H^+ \]

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

\[ Mn^{3+} + 2 H_2O(l) \rightarrow MnO_2(s) + e^- + 4 H^+ \] \[ \text{(1)} \]

The reduction half reaction:

\[ Mn^{3+} \rightarrow Mn^{2+} \]

Add 1 electron to balance the oxidation no.:

\[ Mn^{3+} + e^- \rightarrow Mn^{2+} \] \[ \text{(2)} \]

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

\[ Cl_2 (s) + SO_2 (aq) + H_2O (l) \rightarrow Cl^- (aq) + SO_4^{2-} (aq) \]

The oxidation half reaction:

\[ SO_2 (aq) \rightarrow SO_4^{2-} (aq) \]

Add 2 electrons to balance the oxidation no. :

\[ SO_2 (aq) \rightarrow SO_4^{2-} (aq) + 2 \ e^- \]

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

\[ SO_2 (aq) \rightarrow SO_4^{2-} (aq) + 4 \ H^+ (aq) + 2 \ e^- \]
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:
26. Using the standard electrode potentials given in Table 8.1, predict if the reaction between the following is feasible:

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

(b) \( \text{Ag}^+ \) (aq) and \( \text{Cu} \) (s)
(c) \( \text{Fe}^{3+}_{(aq)} \text{ and } \text{Cu}^{+}_{(s)} \)

(d) \( \text{Ag}_{(s)} \text{ and } \text{Fe}^{3+}_{(aq)} \)

(e) \( \text{Br}_2_{(aq)} \text{ and } \text{Fe}^{2+}_{(aq)} \)

Answer:

(a) \( \text{Fe}^{3+}_{(aq)} \text{ and } \text{I}^{-}_{(aq)} \)

\[ 2 \text{Fe}^{3+}_{(aq)} + 2 \text{I}^{-}_{(aq)} \rightarrow 2 \text{Fe}^{2+}_{(aq)} + \text{I}_2(s) \]

Oxidation half reaction: \( 2 \text{I}^{-}_{(aq)} \rightarrow \text{I}_2(s) + 2e^- \); \( E^\circ = -0.54V \)

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

\[ 2 \text{Fe}^{3+}_{(aq)} + 2 \text{I}^- \rightarrow 2 \text{Fe}^{2+}_{(aq)} + \text{I}_2(s); \]

\( E^\circ = +0.23V \)

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

(b) \( \text{Ag}^{+}_{(aq)} \text{ and } \text{Cu}_{(s)} \)

\[ 2 \text{Ag}^{+}_{(aq)} + \text{Cu}_{(s)} \rightarrow 2 \text{Ag}_{(s)} + \text{Cu}^{2+}_{(aq)} \]

Oxidation half reaction: \( \text{Cu}_{(s)} \rightarrow \text{Cu}^{2+}_{(aq)} + 2e^- \); \( E^\circ = -0.34V \)
Reduction half reaction: \[ \left[ Ag^{+}_{(aq)} + e^- \rightarrow Ag(s) \right] \times 2; \quad E^\circ = +0.80V \]

\[ 2 \; Ag^{+}_{(aq)} + Cu(s) \rightarrow 2 \; Ag(s) + Cu^{2+}; \quad E^\circ = +0.46V \]

\( E^\circ \) 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: \( Cu(s) \rightarrow Cu^{2+}_{(aq)} + 2 \; e^-; \quad E^\circ = -0.34V \)

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

\[ 2 \; Fe^{3+}_{(aq)} + Cu(s) \rightarrow 2 \; Fe^{2+}_{(s)} + Cu^{2+}_{(aq)}; \quad E^\circ = +0.43V \]

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

Reduction half reaction: \[ Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}; \quad E^\circ = +0.77V \]
\[ Ag(s) + Fe^{3+}_{(aq)} \rightarrow Ag^{+}_{(aq)} + Fe^{2+}_{(aq)}; \quad E^\circ = -0.03V \]

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

(e) \( Br_2_{(aq)} \) and \( Fe^{2+}_{(aq)} \)

\[ Br_2(s) + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)} \]

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

Reduction half reaction: \( Br_2_{(aq)} + 2e^- \rightarrow 2Br^-_{(aq)}; \quad E^\circ = +1.09V \)

\[ Br_2(s) + 2Fe^{2+}_{(aq)} \rightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)}; \quad E^\circ = -0.32V \]

\( E^\circ \) 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) $\text{AgNO}_3$ ionizes in aqueous solution to form $\text{Ag}^+$ and $\text{NO}_3^-$ ions.

On electrolysis, either $\text{Ag}^+$ ion or $\text{H}_2\text{O}$ molecule can be decreased at cathode. But the reduction potential of $\text{Ag}^+$ ions is higher than that of $\text{H}_2\text{O}$.

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

$$2\text{H}_2\text{O}_\text{l} + 2e^- \rightarrow \text{H}_2\text{O}_\text{g} + 2\text{OH}^-\text{aq}; E^\circ = -0.83V$$

Therefore, $\text{Ag}^+$ ions are decreased at cathode. Same way, Ag metal or $\text{H}_2\text{O}$ molecules can be oxidized at anode. But the oxidation potential of Ag is greater than that of $\text{H}_2\text{O}$ molecules.

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

$$2\text{H}_2\text{O}_\text{l} \rightarrow \text{O}_2\text{g} + 4\text{H}^+_\text{aq} + 4e^-; E^\circ = -1.23V$$

Hence, Ag metal gets oxidized at anode.

(ii) Pt cannot be oxidized very easily. Therefore, at anode, oxidation of water occurs to liberate $\text{O}_2$. At the cathode, $\text{Ag}^+$ ions are decreased and get deposited.

(iii) $\text{H}_2\text{SO}_4$ ionizes in aqueous solutions to give $\text{H}^+$ and $\text{SO}^{2-}_4$ ions.

$$\text{H}_2\text{SO}_4\text{aq} \rightarrow 2\text{H}_\text{aq}^+ + \text{SO}^{2-}_4\text{aq}$$
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.

\[
Cu^{2+}_{(aq)} + 2 \ e^- \rightarrow Cu_{(aq)} ; \ E^\circ = +0.34V
\]
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
\[ Ag^{+} / Ag = 0.80V \]

\[ Hg^{2+} / Hg = 0.79V \]

\[ Mg^{2+} / Mg = -2.37V \]

\[ Cr^{3+} / Cr = -0.74V \]

**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:

\[ Zn| Zn^{2+}(aq) \parallel Ag^{+}(aq) |Ag \]

(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:

\[ \text{Zn}_\text{(s)} \rightarrow \text{Zn}^{2+}_\text{(aq)} + 2 \text{e}^- \]

Reaction at Ag electrode is shown as:

\[ \text{Ag}^{+}_\text{(aq)} + \text{e}^- \rightarrow \text{Ag}_\text{(s)} \]