# Chapter – 8 (Redox Reactions)

## Exercise Questions:

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

- **a.) $NaH2\underline{P}O4$**
- **b.) $NaH\underline{S}O4$**
- **c.) $H4\underline{P}2O7$**
- **d.) $K2\underline{Mn}O4$**
- **e.) $\underline{Ca}O2$**
- **f.) $Na\underline{B}H4$**
- **g.) $H2\underline{S}2O7$**
- **h.) $KAl(\underline{S}O4)2.12H2O$**

Answer:

(a) **$NaH2PO4$**

Let assume oxidation number of P is x.

We know that,

Oxidation number of Na = +1

Oxidation number of H = +1

Oxidation number of O = -2

Then we have

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

$= 1 + 2 + x - 8 = 0$

$= x - 5 = 0$

$= x = + 5$

Hence the oxidation number of P is +5

**(b) $NaHSO4$**

Let assume oxidation number of S is x.

Oxidation number of Na = +1

Oxidation number of H = +1

Oxidation number of O = -2

Then we have:

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

$= 1 + 1 + x - 8 = 0$

$= x - 6 = 0$

$= x = +6$

Hence the oxidation number of S is +6

**(c) $H_4P_2O_7$**

Let assume oxidation number of P is x.

Oxidation number of H = +1

Oxidation number of O = -2

Then we have:

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

$= 4 + 2x - 14 = 0$

$= 2x - 10 = 0$

$= 2x = +10$

$= x = +5$

Hence, Oxidation number of P is +5

**(d) $K_2MnO_4$**

Let assume oxidation number of Mn is x.

Oxidation number of K = +1

Oxidation number of O = -2

Then we have:

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

$= 2 + x - 8 = 0$

$= x - 6 = 0$

$= x = +6$

Hence, Oxidation number of Mn is +6

**(e) $CaO_2$**

Let assume oxidation number of O is x.

Oxidation number of Ca = +2

Then we have:

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

$= 2 + 2x = 0$

$= 2x = -2$

$= x = -1$

Hence, Oxidation number of O is -1

**(f) NaBH4**

Let assume oxidation number of B is x.

Oxidation number of Na = +1

Oxidation number of H = -1

Then we have:

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

$= 1 + x - 4 = 0$

$= x - 3 = 0$

$= x = +3$

Hence, Oxidation number of B is +3.

**(g) H2S2O7**

Let assume oxidation number of S is x.

Oxidation number of O = -2

Oxidation number of H = +1

Then we have:

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

$= 2 + 2x - 14 = 0$

$= 2x - 12 = 0$

$= x = +6$

Hence, Oxidation number of S is +6.

**(h) KAl(SO4)2.12 H2O**

Let assume oxidation number of S is x.

Oxidation number of K = +1

Oxidation number of Al = +3

Oxidation number of O = -2

Oxidation number of H = +1

Then we have:

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

$= 1 + 3 + 2x - 16 + 24 - 24 = 0$

$= 2x - 12 = 0$

$= 2x = +12$

$= x = +6$

Hence, Oxidation number of S is +6.

**Question: 2 What are the oxidation numbers of the underlined elements in each of the following and how do you rationalise your result?**

- **a.)** $\mathbf{K\underline{I_3}}$
- **b.)** $\mathbf{H_2\underline{S}_4O_6}$
- **c.)** $\mathbf{\underline{Fe}_3O_4}$
- **d.)** $\mathbf{\underline{C}H_3\underline{C}H_2OH}$
- **e.)** $\mathbf{\underline{C}H_3\underline{C}OOH}$

Answer:

**(a) $KI_3$**

Let assume oxidation number of I is x.

In $KI_3$, the oxidation number (O.N.) of K is +1.

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

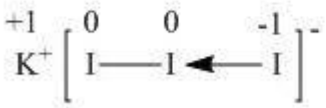
$= +1 + 3x = 0$

$= 3x = -1$

$= x = -1/3$

Hence, the average oxidation number of I is - 1/3

However, O.N. cannot be fractional. Therefore, we will have to consider the structure of $KI_3$ to find the oxidation states. In a $KI_3$ molecule, an atom of iodine forms a coordinate covalent bond with an iodine molecule.



Hence, in a $KI_3$ molecule, the O.N. of the two I atoms forming the $I_2$ molecule is 0, whereas the O.N. of the I atom forming the coordinate bond is –1.

**(b) $H_2S_4O_6$**

Let assume oxidation number of S is x.

The oxidation number (O.N.) of H is +1.

The oxidation number (O.N.) of O is -2.

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

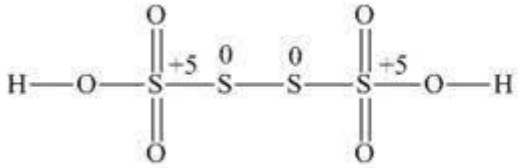
$= 2 + 4x - 12 = 0$

$= 4x - 10 = 0$

$= 4x = +10$

$= x = +10/4$

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



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

**(c) Fe3O4**

Let assume oxidation number of Fe is x.
The oxidation number (O.N.) of O is -2.
$3(x) + 4(-2) = 0$
$= 3x - 8 = 0$
$= 3x = 8$
$= x = 8/3$
However, O.N. cannot be fractional.
Here, one of the three Fe atoms exhibits the O.N. of +2 and the other two Fe atoms exhibit the O.N. of +3.

**(d) CH3CH2OH**

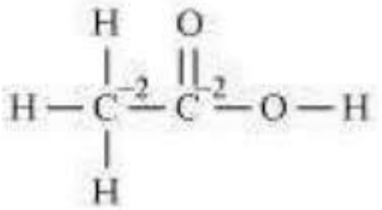
Let assume oxidation number of C is x.
The oxidation number (O.N.) of O is -2.
The oxidation number (O.N.) of H is +1.
$x + 3(+1) + x + 2(+1) + 1(-2) + 1(+1) = 0$
$= x + 3 + x + 2 - 2 + 1 = 0$
$= 2x + 4 = 0$
$= 2x = -4$
$= x = -2$
Hence, the oxidation number of C is -2.

**(e) CH3COOH**

Let assume oxidation number of C is x.
The oxidation number (O.N.) of O is -2.
The oxidation number (O.N.) of H is +1.
$x + 3(+1) + x + (-2) + (-2) + 1(+1) = 0$
$= 2x + 3 - 2 - 2 + 1 = 0$
$= 2x + 0 = 0$

= x = 0

However, 0 is average O.N. of C.



The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number. Thus, C exhibits the oxidation states of +2 and –2 in CH3COOH.

**Question: 3 Justify that the following reactions are redox reactions:**

- **a.) CuO(s) + H2(g) → Cu(s) + H2O(g)**
- **b.) Fe2O3(s) + 3CO(g) → 2Fe(s) + 3CO2(g)**
- **c.) 4BCl3(g) + 3LiAlH4(s) → 2B2H6 + 3LiCl(s) + 3AlCl3(s)**
- **d.) 2K(s) + F2(g) → $2K^{+}F^{-}$9s)**
- **e.) 4NH3(g) + 5O2(g) → 4NO(g) + 6H2O(g)**

Answer:

(a) CuO(s) + H2(g) --> Cu(s) + H2O(g)

Let us write the oxidation number of each element involved in the given reaction as:

+2 -2 0 0 +1 -2

Cu O(s) + H2(g) → Cu(s) + H2 O(g)

Here, the oxidation number of Cu decreases from +2 in CuO to 0 in Cu i.e., CuO is reduced to Cu. Also, the oxidation number of H increases from 0 in H2 to +1 in H2O i.e., H2 is oxidized to H2O. Hence, this reaction is a redox reaction.

(b) Fe2O3(s) + 3CO(g) → 2Fe(s) + 3CO2(g)

Let us write the oxidation number of each element involved in the given reaction as:

+3 -2 +2 -2 0 +4 -2

Fe2 O3(s) + 3C O(g) ⟶ 2Fe(s) + 3C O2(g)

Here, the oxidation number of Fe decreases from +3 in Fe2O3 to 0 in Fe i.e., Fe2O3 is reduced to Fe. On the other hand, the oxidation number of C increases from +2 in CO to +4 in CO2 i.e., CO is oxidized to CO2. Hence, the given reaction is a redox reaction.

(c) 4BCl3(g) + 3LiAlH4(s) ⟶ 2B2H6(g) + 3LiCl(s) + 3 AlCl3 (s)

Let us write the oxidation number of each element involved in the given reaction as:

+3 -1 +1 +3 -1 -3 +1 +1 -1 +3 -1

4B Cl3(g) + 3 Li Al H4(s) ⟶ 2B2 H6(g) + 3Li Cl(s) + 3 Al Cl3 (s)

In this reaction, the oxidation number of B decreases from +3 in BCl3 to –3 in B2H6. i.e., BCl3 is reduced to B2H6. Also, the oxidation number of H increases from –1 in LiAlH4 to +1 in B2H6 i.e., LiAlH4 is oxidized to B2H6. Hence, the given reaction is a redox reaction.

(d) 2K(s) + F2(g) ⟶ 2K+F– (s)

Let us write the oxidation number of each element involved in the given reaction as:

0 0 +1 -1

2K(s) + F2(g) ⟶ 2K+ F– (s)

In this reaction, the oxidation number of K increases from 0 in K to +1 in KF i.e., K is oxidized to KF. On the other hand, the oxidation number of F decreases from 0 in F2 to – 1 in KF i.e., F2 is reduced to KF.

Hence, the above reaction is a redox reaction.

(e) 4 NH3(g) + 5 O2(g) ⟶ 4NO(g) + 6H2O(g)

Let us write the oxidation number of each element involved in the given reaction as:

-3 +1 0 +2 -2 +1 -2

4 N H3(g) + 5 O2(g) ⟶ 4N O(g) + 6H2 O(g)

Here, the oxidation number of N increases from –3 in NH3 to +2 in NO. On the other hand, the oxidation number of O2 decreases from 0 in O2 to –2 in NO and H2O i.e., O2 is reduced. Hence, the given reaction is a redox reaction.

**Question: 4 Fluorine reacts with ice and results in the change:**

**H2O(s) + F2(g) → HF(g) + HOF(g)**

**Justify that this reaction is a redox reaction.**

Answer:

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

+2 -2 0 +1 -1 +1 -2 +1

H2O + F2 → HF + HOF

Here, we have observed that the oxidation number of F increases from 0 in F2 to +1 in HOF. Also, the oxidation number decreases from 0 in F2 to -1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

**Question: 5 Calculate the oxidation number of sulphur, chromium and nitrogen in H2SO5, $Cr2O7^{2-}$, and $NO3^{-}$. Suggest structure of these compounds. Count for the fallacy.**

Answer:

(a) H2SO5

Let assume oxidation number of S is x.

We know that,

Oxidation number of H = +1

Oxidation number of O = -2

Then we have

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

$= 2 + x - 10 = 0$

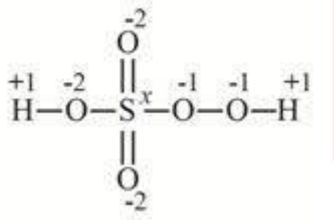
$= x = +8$

However, the O.N. of S cannot be +8. S has six valence electrons. Therefore, the O.N. of S cannot be more than +6. The structure of H2SO5 is shown as follows:

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

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

$= x = +6$



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

(b) Cr2O2-7

Let assume oxidation number of Cr is x.

We know that,

Oxidation number of O = -2

Then we have

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

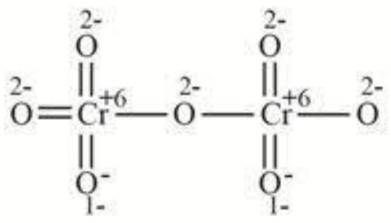
$= 2x - 14 = -2$

$= 2x = +12$

$x = +6$

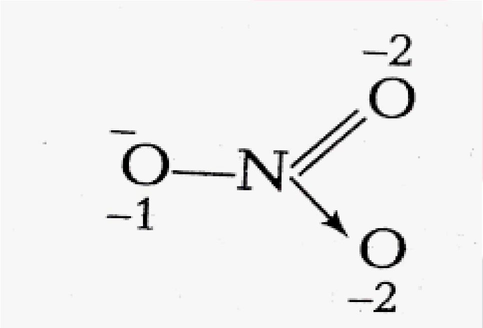
Here, there is no fallacy about the O.N. of Cr in Cr2O72-

The structure of Cr2O72- is shown as follows:



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

(c) NO3-
Let assume oxidation number of N is x.
We know that,
Oxidation number of O = -2
Then we have
1(x) + 3 (-2) = -1
= x - 6 = -1
x = +5
Here, there is no fallacy about the O.N. of N in NO3-
The structure of NO3- is shown as follows:



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

**Question: 6 Write the formulae 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:
$Cr_2O_3$

**Question: 7 Suggest a list of the substances where carbon can exhibit oxidation states from -4 to +4 and nitrogen from -3 to +5.**

Answer:

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

| **Substance** | **O.N. of carbon** |
|---|---|
| $CH_2Cl_2$ | 0 |
| ClC=CCl | +1 |
| HC=CH | -1 |
| $CHCl_3$, CO | +2 |
| $CH_3Cl$ | -2 |
| $Cl_3C$-$CCl_3$ | +3 |
| $H_3C$-$CH_3$ | -3 |
| $CCL_4$, $CO_2$ | +4 |
| $CH_4$ | -4 |

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

| **Substance** | **O.N. of nitrogen** |
|---|---|
| $N_2$ | 0 |
| $N_2O$ | +1 |
| $N_2H_2$ | -1 |
| NO | +2 |
| $N_2H_4$ | -2 |
| $N_2O_3$ | +3 |
| $NH_3$ | -3 |

| | |
|---|---|
| $NO_2$ | +4 |
| $N_2O_5$ | +5 |

**Question: 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 number (O.N.) of S is +4 and the range of the O.N. that S can have is from +6 to -2.

Therefore, $SO_2$ can act as an oxidising as well as a reducing agent.

In hydrogen peroxide ($H_2O_2$), the O.N. of O is -1 and the range of the O.N. that O can have is from 0 to -2. O can sometimes also attain the oxidation numbers +1 and +2. Hence, $H_2O_2$ can act as an oxidising as well as a reducing agent.

In ozone ($O_3$), the O.N. of O is zero and the range of the O.N. that O can have is from 0 to -2. Therefore, the O.N. of O can only decrease in this case. Hence, $O_3$ acts only as an oxidant.

In nitric acid ($HNO_3$), the O.N. of N is +5 and the range of the O.N. that N can have is from +5 to -3. Therefore, the O.N. of N can only decrease in this case. Hence, $HNO_3$ acts only as an oxidant.

**Question: 9 Consider the reactions:**

**a.) $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_5(aq) + 6O_2(g)$**

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

**Why it is more appropriate to write these reaction 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.) The process of photosynthesis involves two steps.

Step 1:

$H_2O$ decomposes to give $H_2$ and $O_2$.

$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$

Step 2:

The $H_2$ produced in step 1 reduces $CO_2$, thereby producing glucose ($C_6H_{12}O_6$) and $H_2O$.

$6CO_2(g) + 12H_2(g) \longrightarrow C_6H_{12}O_6(s) + 6H_2O(l)$

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

$$
\begin{array}{l}
\quad 2H_2O \rightarrow CO_2 + 2H_2] * 6 \\
+ 6CO_2 + 12H_2 \rightarrow C_6H_{12}O_6 + 6H_2O \\
- - - - - - - - - - - - - - - - - - - - - - - - - - - \\
6CO_2 + 12H_2O \rightarrow C_6H_{12}O_6 + 6H_2O
\end{array}
$$

It is more appropriate to write the reaction as given above because water molecules are also produced in the process of photosynthesis.
The path of this reaction can be investigated by using radioactive H2O18 in place of H2O.

(b) O2 is produced from each of the two reactants O3 and H2O2. For this reason, O2 is written twice.
The given reaction involves two steps. First, O3 decomposes to form O2 and O. In the second step, H2O2 reacts with the O produced in the first step, thereby producing H2O and O2.

$$
\begin{array}{l}
O_3(g) \rightarrow O_2(g) + O(g) \\
H_2O_2(l) + O(g) \rightarrow H_2O(l) + O_2(g) \\
- - - - - - - - - - - - - - - - - - - - - - - \\
H_2O_2(l) + O_3(g) \rightarrow H_2O(l) + O_2(g) + O_2(g)
\end{array}
$$

The path of this reaction can be investigated by using H2O18 or O183.

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

Answer:
The oxidation state of Ag in AgF2 is +2. But, +2 is an unstable state of Ag. Therefore, whenever AgF2 is formed, silver readily accepts an electron to form $Ag^+$. This helps to bring the oxidation state of Ag down from +2 to a more stable state +1. As a result, AgF2 acts as a very strong oxidising agent.

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

Answer:
Whenever a reaction between an oxidizing agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing ent is in excess and a compound of higher oxidation state is formed if the oxidizing agent is in excess. Following illustrations justify this.

(i) Oxidizing agent is $F_2$ and reducing agent is $P_4$. When excess $P_4$ reacts with $F_2$, $PF_3$ is produced in which P has +3 oxidation number.

$P_4$( excess ) +$F_2$ → $PF_3$

But if fluorine is in excess, $PF_5$ is formed in which P has oxidation number of +5.

$P_4$ + $F_2$( excess ) → $PF_5$

(ii) Oxidizing agent is oxygen and reducing agent is K. When excess K reacts with oxygen, $K_2O$ is formed in which oxygen has oxidation number of -2.

4K( excess ) + $O_2$ → $2K_2O$

But if oxygen is in excess, then $K_2O_2$ is formed in which O has oxidation number of -1.

2K+ $O_2$( excess ) → $K_2O_2$

(iii) The oxidizing agent is oxygen and the reducing agent is C. When an excess of C reacts with oxygen, CO is formed in which C has +2 oxidation number.

C( excess ) +$O_2$ → CO

When excess of oxygen is used, $CO_2$ is formed in which C has +4 oxidation number.

C+ $O_2$( excess ) → $CO_2$

**Question: 12 How do you count for the following observations?**

**a.) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.**

**b.) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why?**

Answer:

(a) In the manufacture of benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant because of the following reasons.

(i) In a neutral medium, ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii) $KMnO_4$ and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol, $KMnO_4$ and toluene can react at a faster rate.

The balanced redox equation for the reaction in a neutral medium is give as below:

(b) When conc. H2SO4 is added to an inorganic mixture containing bromide, initially HBr is produced. HBr, being a strong reducing agent reduces H2SO4 to SO2 with the evolution of red vapour of bromine.

But, when conc. H2SO4 is added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, being a weak reducing agent, cannot reduce H2SO4 to SO2.

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

- **a.)** $2AgBr(s) + C6H6O2(aq) \rightarrow 2Ag(s) + 2HBr(aq) + C6H4O2(aq)$
- **b.)** $HCHO(l) + 2[Ag(NH3)2]^{+}(aq) + 3OH^{-}(aq) \rightarrow 2Ag(s) + HCOO^{-}(aq) + 4NH3(aq) + 2H2O(l)$
- **c.)** $HCHO(l) + 2Cu^{2+}(aq) + 5OH^{-}(aq) \rightarrow Cu2O(s) + HCOO^{-}(aq) + 3H2O(l)$
- **d.)** $N2H4(l) + 2H2O2(l) \rightarrow N2(g) + 4H2O(l)$
- **e.)** $Pb(s) + PbO2(s) + 2H2SO4(aq) \rightarrow 2PbSO4(s) + 2H2O(l)$

Answer:

a) AgBr is reduced to Ag
C6H6O2 is oxidised to C6H4O2
AgBr is an oxidising agent
C6H6O2 is a reducing agent.

b) [Ag(NH3)+2 reduced to Ag+
HCHO is oxidised to HCOO-
[Ag(NH3)+2 is an oxidising agent.

c) HCHO is oxidised to HCOO-
Cu2+ is reduced to Cu(I) state.

d) N2H4 is reduced to H2O
H2O2 is an oxidising agent
N2H4 is reducing agent.

e) Pb has been oxidised to PbSO4
PbO2 is reduced to PbSO4
PbO2 is an oxidising agent
Pb is a reducing agent

**Question: 14 Consider the reactions:**

**$2S2O3^{2-} + I2(s) \rightarrow S4O6^{2-}(aq) + 2I^{-}(aq)$**

**$S2O3^{2-}(aq) + 2Br2(l) + 5H2O(l) \rightarrow 2SO4^{2-}(aq) + 4Br^{-}(aq) + 10H^{+}(aq)$**

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

Answer:

+2 -2 0 2.5 -2

$2S2O3^{2-} + I2(s) \rightarrow S4O6^{2-}(aq) + 2I^{-}(aq)$

+2 -2 0 +6 -2

$S2O3^{2-}(aq) + 2Br2(l) + 5H2O(l) \rightarrow 2SO4^{2-}(aq) + 4Br^{-}(aq) + 10H^{+}(aq)$

Bromine is a stronger oxidising agent when compared to I2. It oxidises the S of $S2O3^{3-}$ to a higher oxidation state +6 in $SO4^{2-}$. While I2 oxidises S of $S2O3^{2-}$ to a lower oxidation state 2.5 in $S4O6^{2-}$. That's why same reductant, thiosulphate react differently with bromine and iodine.

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

Answer:

F2 can oxidize Cl- to Cl2, Br- to Br2, and I- to I2 as:

$F2(aq) + S2cl-(s) \rightarrow 2F-(aq) + Cl (g)$

$F2(aq) + 2Br-(aq) \rightarrow 2F-(aq) + Br 2(l)$

$F2(aq) + 2l-(aq) \rightarrow 2F-(aq) + 12(s)$

On the other hand, Cl2, Br2, and I2 cannot oxidize F- to F2. The oxidizing power of halogens increases in the order of I2 < Br2 < Cl2 < F2. Hence, fluorine is the best oxidant among halogens.

HI and HBr can reduce H2SO4 to SO2, but HCl and HF cannot. Therefore, HI and HBr are stronger reductants than HCl and HF.

$2HI + H2SO4 \rightarrow 12 + SO2 + 2H2O$

$2HBr + H2SO4 \rightarrow Br2 + SO2 + 2H2O$

Again, I- can reduce Cu2+ to Cu+, but Br- cannot.

$4l-(aq) + 2Cu2+(aq) \rightarrow Cu2l2(s) + 12(aq)$

Hence, hydroiodic acid is the best reductant among hydrohalic compounds.

Thus, the reducing power of hydrohalic acids increases in the order of HF < HCl < HBr < HI.

**Question: 16 Why does the following reactions occur?**

**$XeO_6^{4-}(aq) + 2F^{-}(aq) + 6H^{+}(aq) \rightarrow XeO3(g) + F2(g) + 3H2o(l)$**

**What conclusion about the compound Na4XeO6 (of which XeO$_6^{4-}$ is a part) can be drawn from the reaction.**

Answer:

The given reaction occurs because XeO6$^{4-}$ oxidises F$^-$ and F$^-$ reduces XeO6$^{4-}$.

+8 -2 +6 0

XeO6$^{4-}$(aq) + 2F$^-$(aq) + 6H$^+$(aq) → XeO3(g) + F2(g) + 3H2O(l)

In this reaction, the oxidation number (O.N.) of Xe decreases from +8 in XeO6$^{4-}$ to +6 in XeO3 and the O.N. of F increases from -1 in F$^-$ to O in F2.

Hence, we can conclude that Na4XeO6 is a stronger oxidising agent than F$^-$.

**Question: 17 Consider the reactions:**

- **a.) H3PO2(aq) + 4AgNO3(aq) + 2H2O(l) → H3PO4(aq) + 4Ag(s) + 4HNO3(aq)**
- **b.) H3PO2(aq) + 2CuSO4(aq) + 2H2O(l) → H3PO4(aq) + 2Cu(s) + H2SO4(aq)**
- **c.) C6H5CHO(l) + 2[Ag(NH3)2]$^+$(aq) + 3OH$^-$(aq) → C6H5COO$^-$(aq) + 2Ag(s) + 4NH3(aq)**
- **d.) C6H5CHO(l) + 2Cu$^{2+}$(aq) + 5OH(aq) → no change observed.**

**What inference do you draw about the behaviour of Ag$^+$ and Cu$^{2+}$ from these reactions?**

Answer:

Ag$^+$ and Cu$^{2+}$ act as oxidising agents in reactions (a) and (b) respectively. In reaction (c), Ag$^+$ oxidises C6H5CHO to C6H5COO$^-$, but in reaction (d), Cu$^{2+}$ cannot oxidises C6H5CHO.

Hence, we can say that Ag$^+$ is a stronger oxidising agent than Cu$^{2+}$.

**Question: 18 Balance the following redox reactions by ion-electron method:**

- **a.) MnO4$^-$(aq) + I$^-$(aq) → MnO2(s) + I2(s) (in basic medium)**
- **b.) MnO4$^-$(aq) + So2(g) → Mn$^{2+}$(aq) + HSO4$^-$(aq) (in acidic solution)**
- **c.) H2O2(aq) + Fe$^{2+}$(aq) → Fe$^{3+}$(aq) + H2O(l) (in acidic medium)**
- **d.) Cr2O7$^{2-}$ + SO2(g) → Cr$^{3+}$(aq) + SO4$^{2-}$(aq) (in acidic solution)**

Answer:

Step 1:

The two half reactions involved in the given reaction are:

-1 0

Oxidation half reaction: l (aq) → l2(s)

+7 +4

Reduction half reaction: Mn O-4(aq) ⟶ MnO2(aq)

Step 2:

Balancing I in the oxidation half reaction, we have:

2l-(aq) → l2(s)

Now, to balance the charge, we add 2 e- to the RHS of the reaction.

2l-(aq) → l2(s) + 2e-

Step 3 :

In the reduction half reaction, the oxidation state of Mn has reduced from +7 to +4. Thus, 3 electrons are added to the LHS of the reaction.

MnO-4(aq) + 3e- → MnO2(aq)

Now, to balance the charge, we add 4 OH- ions to the RHS of the reaction as the reaction is taking place in a basic medium.

MnO-4(aq) + 3e- → MnO2(aq) + 4OH-

Step 4:

In this equation, there are 6 O atoms on the RHS and 4 O atoms on the LHS. Therefore, two water molecules are added to the LHS.

MnO-4(aq) + 2H2O + 3e- → MnO2(aq) + 4OH-

Step 5:

Equalising the number of electrons by multiplying the oxidation half reaction by 3 and the reduction half reaction by 2, we have:

6l-(aq) → 3l2(s) + 2e-

2MnO-4(aq) + 4H2O + 6e- → 2MnO2(s) + 8OH-(aq)

Step 6:

Adding the two half reactions, we have the net balanced redox reaction as:

6l-(aq) + 2MnO-4(aq) + 4H2O(l) → 3l2(s) + 2MnO2(s) + 8OH-(aq)

a.) Following the steps as in part (a), we have the oxidation half reaction as:

SO2(g) + 2H2O(l) → HSO-4(aq) + 3H+(aq) + 2e-(aq)

And the reduction half reaction as:

MnO-4(aq) + 8H+(aq) + 5e- → Mn2+(aq) + 4H2O(l)

Multiplying the oxidation half reaction by 5 and the reduction half reaction by 2, and then by adding them, we have the net balanced redox reaction as:

$2MnO-4(aq) + 5SO2(g) + 2H2O(l) + H+(aq) \rightarrow Mn2+(aq) + HSO-4(aq)$

(a) Following the steps as in part (a), we have the oxidation half reaction as:

$Fe2+(aq) \rightarrow Fe3+(aq) + e-$

And the reduction half reaction as:

$H2O2(aq) + 2H+(aq) + 2e- \rightarrow 2H2O(l)$

Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:

$H2O2(aq) + 2Fe2+(aq) + 2H+(aq) \rightarrow 2Fe3+(aq) + 2H2O(l)$

(b) Following the steps as in part (a), we have the oxidation half reaction as:

$SO2(g) + 2H2O(l) \rightarrow SO2-4(aq) + 4H+ (aq) + 2e-$

And the reduction half reaction as:

$Cr2O2-7(aq) + 14H+(aq) + 6e- \rightarrow 2Cr3+(aq) + 3SO2-4(aq) + H2O(l)$

Multiplying the oxidation half reaction by 3 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:

$Cr2O2-7(aq) + 3SO2(g) + 2H+(aq) \rightarrow 2Cr3+(aq) + 3SO2-4(aq) + H2O(l)$

**Question: 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.) $P4(s) + OH^-(aq) \rightarrow PH3(g) + HPO2^-(aq)$**

**b.) $N2H4(l) + ClO3^-(aq) \rightarrow NO(g) + Cl^{-(aq)}$**

**c.) $Cl2O7(g) + H2O2(aq) \rightarrow ClO2^-(aq) + O2(g) + H^+(aq)$**

Answer:

(a) The O.N. (oxidation number) of P decreases from 0 in P4 to -3 in PH3 and increases from 0 in P4 to + 2 in HPO-2. Hence, P4 acts both as an oxidizing agent and a reducing agent in this reaction.

Ion-electron method:

The oxidation half equation is:

$P4(s) \rightarrow H2PO-(aq)$

The P atom is balanced as:

$\overset{0}{P4}(s) \longrightarrow 4\overset{2+}{HPO2}-(aq)$

The O.N. is balanced by adding 4 electrons as:

P4(s) → 4H2PO-(aq) + 4e-

The charge is balanced by adding 8OH- as:

P4(s) + 8OH - (aq) → 4H2PO-2(aq)

The O and H atoms are already balanced. The reduction half equation is:

P4(s) → PH3(g)

The P atom is balanced as

0 -3

P4(s) → 4 PH3(g)

The O.N. is balanced by adding 12 electrons as:

P4(s) + 12e- → 4 PH3(g)

The charge is balanced by adding 12OH- as:

P4(s) + 12e- → 4 PH3(g) + 12OH-(aq) .....(i)

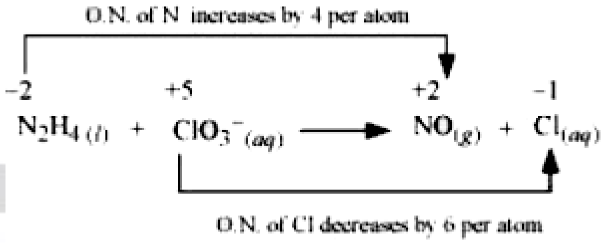
The O and H atoms are balanced by adding 12H2O as:

P4(s) + 12H2O(l) + 12e- → 4 PH3(g) + 12OH-(aq) -- (ii)

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

P4(s) + 3OH-(aq) + 3H2O → PH3 + 3H2PO-2(aq)

**(b)**



The oxidation number of N increases from -2 in N2H4 to +2 in NO and the oxidation number of Cl decreases from + 5 in ClO-3 to -1 in Cl-. Hence, in this reaction, N2H4 is the reducing agent and ClO-3 is the oxidizing agent. Ion-electron method:

The oxidation half equation is:

-2 +2

N2 H4(l) → N O(g)

The N atoms are balanced as:

N2H4(l) → 2NO(g)
The oxidation number is balanced by adding 8 electrons as:
N2H4(l) → 2NO(g) + 8e-
The charge is balanced by adding 8 OH-ions as:
N2H4(l) + 8OH-(aq) ⟶ 2NO(g) + 8e-
The O atoms are balanced by adding 6H2O as:
N2H4(l) + 8OH-(aq) ⟶ 2NO(g) + 6H2O(l) + 8e- .... (i)
The reduction half equation is:
+5 -1

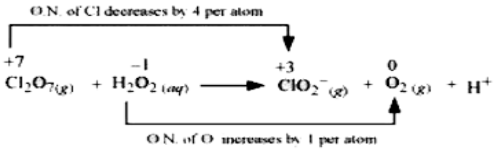
CIO-3(aq) ⟶ Cl-(aq)

The oxidation number is balanced by adding 6 electrons as
CIO-3(aq) + 6e- ⟶ Cl-(aq)
The charge is balanced by adding 6OH- ions as:
CIO-3(aq) + 6e- ⟶ Cl-(aq) + 6OH-(aq)
The O atoms are balanced by adding 3H2O as:
CIO-3(aq) + 3H2O(l) + 6e- ⟶ Cl-(aq) + 6OH-(aq) .... (ii)

The balanced equation can be obtained by multiplying equation (i) with 3 and equation (ii) with 4 and then adding them as:
3N2H4(l) + 4CIO-3(aq) ⟶ 6NO(g) + 4Cl-(aq) + 6H2O(l)
Oxidation number method:
Total decrease in oxidation number of N = 2 × 4 = 8
Total increase in oxidation number of Cl = 1 × 6 = 6
On multiplying N2H4 with 3 and CIO-3 with 4 to balance the increase and decrease in O.N., we get:
3N2H4(l) + 4CIO-3(aq) ⟶ NO(g) + Cl-(aq)
The N and Cl atoms are balanced as:
3N2H4(l) + 4CIO-3(aq) ⟶ 6NO(g) + 4Cl-(aq)

The O atoms are balanced by adding 6H2O as:
3N2H4(l) + 4CIO-3(aq) ⟶ 6NO(g) + 4Cl-(aq) + 6H2O(l)
This is the required balanced equation.

(c)



The oxidation number of Cl decreases from + 7 in Cl2O7 to + 3 in ClO-2and the oxidation number of O increases from -1 in H2O2 to zero in O2. Hence, in this reaction, Cl2O7 is the oxidizing agent and H2O2 is the reducing agent.

Ion-electron method:

The oxidation half equation is:

$\overset{-1}{\mathrm{H2O2}}\mathrm{(aq)} \rightarrow \overset{0}{\mathrm{O2}}\mathrm{(g)}$

The oxidation number is balanced by adding 2 electrons as:

$\mathrm{H2O2(aq)} \rightarrow \mathrm{O2(g)} + \mathrm{2e-}$

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

$\mathrm{H2O2(aq)} + \mathrm{2OH\text{-}(aq)} \longrightarrow \mathrm{O2(g)} + \mathrm{2e-}$

The oxygen atoms are balanced by adding 2H2O as:

$\mathrm{H2O2(aq)} + \mathrm{2OH\text{-}(aq)} \longrightarrow \mathrm{O2(g)} + \mathrm{2H2O(l)} + \mathrm{2e-}$ ... (i)

The reduction half equation is:

$\overset{+7}{\mathrm{Cl2O7}}\mathrm{(g)} \rightarrow \overset{+3}{\mathrm{ClO\text{-}2}}\mathrm{(g)}$

The Cl atoms are balanced as:

$\mathrm{Cl2O7(g)} \rightarrow \mathrm{2ClO\text{-}2(g)}$

The oxidation number is balanced by adding 8 electrons as:

$\mathrm{Cl2O7(g)} + \mathrm{8e-} \rightarrow \mathrm{2ClO\text{-}2(g)}$

The charge is balanced by adding 6OH- as:

$\mathrm{Cl2O7(g)} + \mathrm{8e-} \rightarrow \mathrm{2ClO\text{-}2(g)} + \mathrm{6OH\text{-}\ (aq)}$

The oxygen atoms are balanced by adding 3H2O as:

$\mathrm{Cl2O7(g)} + \mathrm{3H2O(l)} + \mathrm{8e-} \rightarrow \mathrm{2ClO\text{-}2(g)} + \mathrm{6OH\text{-}\ (aq)}$ .... (ii)

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

$\mathrm{Cl2O7(g)} + \mathrm{4H2O2(aq)} + \mathrm{2OH\text{-}\ (aq)} \rightarrow \mathrm{2ClO\text{-}2(aq)} + \mathrm{4O2(g)} + \mathrm{5H2O(l)}$

Oxidation number method:

Total decrease in oxidation number of Cl2O7 = $4 \times 2 = 8$

Total increase in oxidation number of H2O2 = $2 \times 1 = 2$

By multiplying H2O2 and O2 with 4 to balance the increase and decrease in the oxidation number, we get:

$Cl2O7(g) + 4H2O2(aq) \rightarrow CIO\text{-}2(aq) + 4O2(g)$

The Cl atoms are balanced as:

$Cl2O7(g) + 4H2O2(aq) \rightarrow 2CIO\text{-}2(aq) + 4O2(g)$

The O atoms are balanced by adding 3H2O as:

$Cl2O7(g) + 4H2O2(aq) \rightarrow 2CIO\text{-}2(aq) + 4O2(g) + 3H2O(l)$

The H atoms are balanced by adding 2OH- and 2H2O as:

$Cl2O7(g) + 4H2O2(aq) + 2OH\text{-}(aq) \rightarrow 2CIO\text{-}2(aq) + 4O2(g) + 5H2O(l)$

This is the required balanced equation.

**Question: 20 What sorts of informations can you draw from the following reaction?**

**$(CN)2(g) + 2OH^{-}(aq) \rightarrow CN^{-}(aq) + CNO^{-}(aq) + H2O(l)$.**

Answer:

The oxidation numbers of carbon in (CN)2 , CN– and CNO– are +3, +2 and +4 respectively. These are obtained as shown below:

Let the oxidation number of C be x.

(CN)2

$2(x – 3) = 0$

$x = 3$

CN– $x – 3 = –1$

$x = 2$

CNO–

$x – 3 – 2 = –1$

$x = 4$

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

**Question: 21 The $Mn^{3+}$ ion is unstable in solution and undergoes disproportionation to give $Mn^{2+}$, MnO2, and $H^{+}$ ion. Write a balanced ionic equation for the reaction.**

Answer:

Oxidation Reaction

Mn+3 --> MnO2

Balance by adding electrons

Mn+3 --> MnO2 + e-

Mn+3 + H2O --> MnO2 + 4H+ + e-

Reduction Reaction

Mn+3 --> Mn+2

Balance by adding electrons

Mn+3 + e- --> Mn+2

Adding equations 1 and 2

2Mn+3 + 2H2O --> MnO2 + Mn+2 + 4H+

**Question: 22 Consider the elements:**

**Cs, Ne, I and F**

- **a.) Identify the element that exhibits only negative oxidation state.**
- **b.) Identify the element that exhibit only positive oxidation state.**
- **c.) Identify the element that exhibits both positive and negative oxidation states.**
- **d.) Identify the element which exhibits neither the negative nor does the positive oxidation state.**

Answer:

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

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

Answer:

The given redox reaction can be represented as:

Cl2(s) + SO2(aq) + H2O(l) → Cl-(aq) + SO2-4(aq)

The oxidation half reaction is:

S + 4O2(aq) → S+6O2-4(aq)

The oxidation number is balanced by adding two electrons as:

SO2(aq) → SO2-4(aq) + 2e-

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

SO2(aq) → SO2-4(aq) + 4H+(aq) + 2e-

The O atoms and H+ ions are balanced by adding 2H2O molecules as:

SO2(aq) + 2 H2O(l) → SO2-4(aq) + 4H+(aq) + 2e- --------(i)

The reduction half reaction is: The chlorine atoms are balanced as:

Cl2(s) → Cl-(aq)

The oxidation number is balanced by adding electrons

Cl2(s) + 2e- → 2 Cl-(aq) ----------(ii)

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

Cl2(s) + SO2(aq) + 2 H2O(l) → 2 Cl-(aq) + SO2-4(aq) + 4H+(aq)

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

  **a.) Select the possible non – metals that can show disproportionation reaction.**
  **b.) Select three metals that can show disproportionation reaction.**

Answer:

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

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

b.) Mn, Cu, and Ga can show disproportionation reactions as these elements can exists in three or more oxidation states.

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

Answer:

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

$4NH3(g) + 5O2(g) \longrightarrow 4NO(g) + 6\ H2O(g)$

4x17g 2x32g 4x30g 6x18g

= 68g =160g =120g 108g

Thus, 68 g of NH3 reacts with 160 g of O2.

Therefore, 10g of NH3 reacts with 160x10 / 68g of O2, or 23.53 g of O2.

But the available amount of O2 is 20 g.

Therefore, O2 is the limiting reagent (we have considered the amount of O2 to calculate the weight of nitric oxide obtained in the reaction).

Now, 160 g of O2 gives 120g of NO.

Therefore, 20 g of O2 gives120x20 / 160 g of N, or 15 g of NO.

Hence, a maximum of 15 g of nitric oxide can be obtained.

**Question: 26 Using the standard electrode potentials given in the table 8.1, predict if the reaction between the following is feasible:**

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

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

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

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

**e.) Br2(aq) and $Fe^{2+}$(aq)**

Answer:

$$\begin{array}{ll}\text{Oxidation half equation:} & 2I^-_{(aq)} \rightarrow I_{2(s)} + 2e^-;\ E^\circ = -0.54\,V \\ \text{Reduction half equation:} & [Fe^{3+}_{(aq)} + e^-] \rightarrow Fe^{2+}_{(aq)} \times 2;\ E^\circ = +0.77\,V \\ \hline & 2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + I_{2(s)};\ E^\circ = +0.23\,V\end{array}$$

i.)

Since Eº for the overall reaction is positive, the reaction between Fe3+(aq) and I - (aq) is feasible.

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

$$\begin{array}{lll}\text{Oxidation half equation :} & Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^- & ;\ E^\circ = -0.34\,V \\ \text{Reduction half equation :} & [Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}] \times 2 & ;\ E^\circ = +0.80\,V \\ \hline & 2Ag^+_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+};\ E^\circ = +0.46\,V & \end{array}$$

ii.)

Since Eº for the overall reaction is positive, the reaction between Ag+(aq) and Cu(s) is feasible.

$$\begin{array}{lll}\text{Oxidation half equation :} & Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^- & ;\ E^\circ = -0.34\,V \\ \text{Reduction half equation :} & [Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(s)}] \times 2 & ;\ E^\circ = +0.77\,V \\ \hline & 2Fe^{3+}_{(aq)} + Cu_{(s)} \longrightarrow 2Fe^{2+}_{(s)} + Cu^{2+}_{(aq)};\ E^\circ = +0.43\,V & \end{array}$$

iii.)

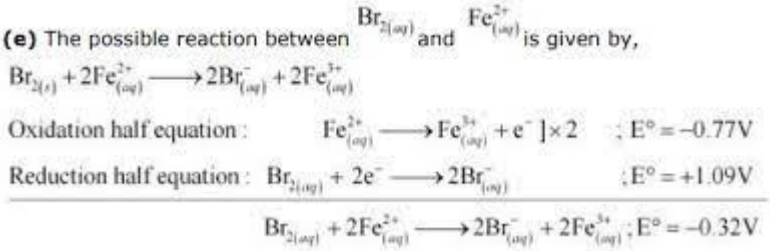
Since Eº for the overall reaction is negative, the reaction between Fe3+(aq) and Br - (aq) is not feasible.

$$\begin{array}{lll}\text{(iv)} & Ag_{(s)} \longrightarrow Ag^+_{(aq)} + e^- & ;\ E^\circ = -0.80\ V \\ & Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)} & ;\ E^\circ = +0.77\ V \\ \hline & Ag_{(s)} + Fe^{3+}_{(aq)} \longrightarrow Ag^+_{(aq)} + Fe^{2+}_{(aq)} & ;\ E^\circ = -0.03\ V\end{array}$$

iv.)

Since Eº for the overall reaction is negative, the reaction between Ag(s) and Fe3+(aq) is not feasible.

v.)



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

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

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

Reduction half equation : $Br_{2(aq)} + 2e^{-} \longrightarrow 2Br^{-}_{(aq)}$ ; $E° = +1.09V$

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

Since for the overall reaction is positive, the reaction between Br2(aq) and Fe2+(aq) is feasible.

**Question: 27 Predict the products of electrolysis in each of the following:**

**a.) An aqueous solution of AgNO3 with silver electrodes.**
**b.) An aqueous solution AgNO3 with platinum electrodes.**
**c.) A dilute solution of H2SO4 with platinum electrodes.**
**d.) Ab aqueous solution of CuCl2 with platinum electrodes.**

Answer:

(i) At cathode: The following reduction reactions compete to take place at the cathode.

$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$ ; $E^{0} = 0.80V$

$H^{+} + e^{-} \rightarrow \frac{1}{2} H2(g)$; $E^{0} = 0.00V$

The reaction with a higher value of E° takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

The Ag anode is attacked by NO3- ions. Therefore, the silver electrode at the anode dissolves in the solution to form Ag+.

(ii) At cathode:

The following reduction reactions compete to take place at the cathode.

$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$ ; $E^{0} = 0.80V$

$H^{+} + e^{-} \rightarrow \frac{1}{2} H2(g)$; $E^{0} = 0.00V$

The reaction with a higher value of E° takes place at the cathode. Therefore, deposition of silver will take place at the cathode.

At anode:

Since Pt electrodes are inert, the anode is not attacked by NO3- ions. Therefore, OH - or NO3- ions can be oxidized at the anode. But OH - ions having a lower discharge potential and get preference and decompose to liberate O2.

OH- ⟶ OH + e-

$4OH^- \longrightarrow 2H_2O + O_2$

(iii) At the cathode, the following reduction reaction occurs to produce $H_2$ gas.

$H^+ (aq) + e^- \longrightarrow \frac{1}{2} H_2(g)$

At the anode, the following processes are possible.

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-; E^0 = 1.23V$

$2SO_4^{2-}(aq) \rightarrow S_2O_6^{2-}(aq) + 2e^-; \ E^0 = 1.96$

For dilute sulphuric acid, reaction (i) is preferred to produce $O_2$ gas. But for concentrated sulphuric acid, reaction (ii) occurs.

(iv) At cathode:

The following reduction reactions compete to take place at the cathode.

$Cu^{2+} (aq) + 2e^- \rightarrow Cu(s); E^0 = 0.34V$

$H^+ + e^- \rightarrow \frac{1}{2} H_2(g); \ E^0 = 0.00V$

The reaction with a higher value of E° takes place at the cathode. Therefore, deposition of copper will take place at the cathode.

At anode:

The following oxidation reactions are possible at the anode.

$Cl^-(aq) \rightarrow \frac{1}{2} Cl_2(g) + e^-; E^0 = 1.36V$

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-; E^0 = 1.23V$

At the anode, the reaction with a lower value of E° is preferred. But due to the over-potential of oxygen, $Cl^-$ gets oxidized at the anode to produce $Cl_2$ gas.

**Question: 28 Arrange the following metals in the order in which they displace each other from the solution of their salts.**

**Al, Cu, Fe, Mg and Zn.**

Answer:

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

The order of the increasing reducing power of the given metals is Cu < Fe < Zn < Al < Mg.

Hence, we can say that Mg can displace Al from its salt solution, but Al cannot displace Mg. Thus, the order in which the given metals displace each other from the solution of their salts is given below:

Mg > Al > Zn > Fe > Cu.

**Question: 29 Given the standard electrodes potential,**

**$K^+/K = -2.93V$, $Ag^+/Ag = 0.80V$**

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

**$Mg^{2+}/Mg = -2.37V$, $Cr^{3+}/Cr = -0.74V$**

**Arrange these metals in their increasing order of reducing power.**

Answer:

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

**Question: 30 Depict the galvanic cell in which the reaction**

**$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place, further show :**

- **a.) Which of the electrode is negatively charged.**
- **b.) The carriers of the current in the cell and**
- **c.) Individuals reaction at each electrode.**

Answer:

The galvanic cell in which the given reaction takes place is depicted as:

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

(i) Zn electrode (anode) is negatively charged.

(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.

(iii) The reaction taking place at the anode is given by,

$Zn(s) \rightarrow Zn2+(aq) + 2e-$

The reaction taking place at the cathode is given by,

$Ag+(aq) + e- \rightarrow Ag(s)$