CBSE NCERT Solutions for Class 11 Chemistry Chapter 8

Back of Chapter Questions

1. Assign O.N. to the underlined elements in each of the following species:

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

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

(c) $H_4\underline{P_2}O_7$

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

(e) $Ca\underline{O_2}$

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

(g) $H_2\underline{S_2}O_7$

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

Solution:

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

Let the O.N. of P = x.

We know that,

The O.N. of Na $= +1$

The O.N. of H $= +1$

The O.N. of O $= -2$

Then, we have

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

$1 + 2 + x - 8 = 0$

$\Rightarrow x = +5$

Hence, the O.N. of P is $+5$

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

Let the O.N. of S = x.

We know that,

The O.N. of Na $= +1$

The O.N. of H $= +1$

The O.N. of $\mathrm{O} = -2$

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

$\Rightarrow 1 + 1 + \mathrm{x} - 8 = 0$

$\Rightarrow \mathrm{x} = +6$

Hence, the O.N. of S is $+6$.

(c) $\mathrm{H_4\underline{P}_2O_7}$

Let the O.N. of P = x.

We know that,

The O.N. of H $= +1$

The O.N. of $\mathrm{O} = -2$

Then, we have

$4(+1) + 2(\mathrm{x}) + 7(-2) = 0$

$\Rightarrow 4 + 2\mathrm{x} - 14 = 0$

$\Rightarrow 2\mathrm{x} = +10$

$\Rightarrow \mathrm{x} = +5$

Hence, the O.N. of P is $+5$

(d) $\mathrm{K_2\underline{Mn}O_4}$

Let the O.N. of Mn = x.

We know that,

The O.N. of K $= +1$

The O.N. of $\mathrm{O} = -2$

Then, we have

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

$\Rightarrow \mathrm{x} = +6$

Hence, the O.N. of Mn is $+6$

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

Let the O.N. of O = x.

We know that,

The O.N. of Ca $= +2$

Then, we have

$(+2) + 2(x) + 0$

$\Rightarrow 2 + 2x = 0$

$\Rightarrow x = -1$

Hence, the O.N. of O is -1

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

Let the O.N. of B = x.

We know that,

The O.N. of Na $= +1$

The O.N. of H $= +1$

Then, we have

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

$\Rightarrow 1 + x - 4 = 0$

$\Rightarrow x = -3$

Hence, the O.N. of B is $+3$

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

Let the O.N. of S = x.

We know that,

The O.N. of H $= +1$

The O.N. of O $= -2$

Then, we have

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

$\Rightarrow 2 + 2x - 14 = 0$

$\Rightarrow 2x = 12$

Hence, the O.N. of S is $+6$

(h) $\mathrm{KAl}\left(\underline{\mathrm{S}}\mathrm{O}_4\right)_2 . 12\ \mathrm{H_2O}$

Let the O.N. of S = x.

We know that,

The O.N. of K$= +1$

The O.N. of Al $= +3$

The O.N. of $\mathrm{O} = -2$

The O.N. of H $= +1$

Then, we have

$1(+1) + 1(+3) + 2\mathrm{x} + 8(-2) + 24(+1) + 12(-2) = 0$

$\Rightarrow 1 + 3 + 2\mathrm{x} - 16 + 24 - 24 = 0$

$\Rightarrow \mathrm{x} = +6$

Hence, the O.N. of S is $+6$

Or

We can ignore the water molecule as it is a neutral molecule. Then, the sum of the O.N.s of all atoms of the water molecule may be taken as zero. Therefore, after ignoring the water molecule, we have

Let the O.N. of S = x.

We know that,

The O.N. of K$= +1$

The O.N. of $\mathrm{O} = -2$

The O.N. of $Al = +3$

The O.N. of $\mathrm{H_2O} = 0$

Then, we have

$1(+1) + 1(+3) + 2\mathrm{x} + 8(-2) = 0$

$\Rightarrow 1 + 3 + 2\mathrm{x} - 16 = 0$

$\Rightarrow \mathrm{x} = +6$

Hence, the O.N. of S is $+6$

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

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

Solution:

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

Let the O.N. of S = x.

We know that,

the O.N. $(\mathrm{O.N.})$ of K = $+1$.

Hence, the average O.N. of I

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

$$\Rightarrow 1 + 3x = 0$$

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

However, $\mathrm{O.N.}$ 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.

$$\overset{+1}{K^+}\left[\overset{0}{I}-\overset{0}{I}\leftarrow\overset{-1}{I}\right]^-$$

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

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

Let the O.N. of S = x.

We know that,

The O.N. of H$= +1$

The O.N. of $O = -2$

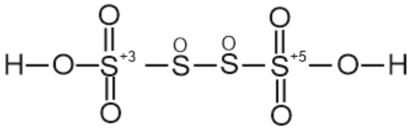
Now, $2(+1) + 4(x) + 6(-2) = 0$

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

$\Rightarrow 4x = 10$

$\Rightarrow x = +2.5$

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



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

(c) Fe_3O_4

Let the O.N. of Fe = x.

We know that,

The O.N. of $O = -2$

Now, $3(x) + 4(-2) = 0$

$\Rightarrow 3x - 8 = 0$

$\Rightarrow 3x = 8$

$\Rightarrow x = +2.5$

the O.N. of Fe is found to be $+2.5$. However, O.N. cannot be fractional.

Fe_3O_4 *is* mixed oxide is has FeO and Fe_2O_3 ,

the oxidation state of Fe in FeO:

x+1(-2)=0

x=2

the oxidation state of Fe in Fe_2O_3:

2x+3(-2)=0

x=3

the O. N. of Fe is found to be 2 and 3 in FeO and Fe_2O_3 respectively.

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

Let the O.N. of C= x.

We know that,

The O.N. of H= $+1$

The O.N. of $O = -2$

Now,

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

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

$\Rightarrow x = -2$

Hence, the O. N. of C is -2.

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

Let the O.N. of C = x.

We know that,

The O.N. of H= $+1$

The O.N. of $O = -2$

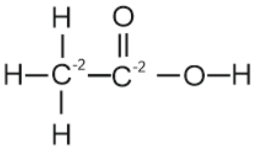
Now, $2(x) + 4(1) + 2(-2) = 0$

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

$\Rightarrow x = 0$

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 O.N..

Thus, C exhibits the oxidation states of $+2$ and -2 in CH_3COOH.



3. Justify that the following reactions are redox reactions:

(a) $\mathrm{CuO(s)\ + H_2(g) + Cu(s) + H_2O(g)}$

(b) $\mathrm{Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s)\ +\ 3CO_2(g)}$

(c) $\mathrm{4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3(s)}$

(d) $\mathrm{2K(s) + F_2(g) \rightarrow 2K^+F^-(s)}$

(e) $\mathrm{4\ NH_3 + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)}$

Solution:

(a) $\mathrm{CuO(s)\ + H_2(g) + Cu(s) + H_2O(g)}$

Let us write the O.N. of each element involved in the given reaction as:

$$\overset{+2}{\mathrm{Cu}}\,\overset{-2}{\mathrm{O}}\,\mathrm{(s)} + \overset{0}{\mathrm{H_2}}\mathrm{(g)} \rightarrow \overset{0}{\mathrm{Cu}}\mathrm{(s)} + \overset{+1}{\mathrm{H_2}}\,\overset{-2}{\mathrm{O}}\,\mathrm{(g)}$$

Here,

$2e^- + Cu^{+2} \rightarrow Cu$

the O.N. of Cu decreases from +2 in CuO to 0 in Cu, i.e., CuO is reduced to Cu. Or in the above reaction addition of electron is taking place so it is a reduction.

$$\overset{0}{\mathrm{H_2}}\mathrm{(g)} \rightarrow \overset{0}{\mathrm{Cu}}\mathrm{(s)} + \overset{+1}{\mathrm{H_2}}\,\overset{-2}{\mathrm{O}}$$

Also, the O.N. of H increases from *zero* in $\mathrm{H_2}$ to +1 in $\mathrm{H_2O}$ i.e., $\mathrm{H_2}$ is oxidized to $\mathrm{H_2O}$.

In the above reaction oxidation and reduction, both is taking place simultaneously Hence, this reaction is a redox reaction.

(b) $\mathrm{Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s)\ +\ 3CO_2(g)}$

Let us write the O.N. of each element in the given reaction as:

$$\overset{+3}{\mathrm{Fe_2}}\,\overset{-2}{\mathrm{O_3}}\mathrm{(s)} + 3\,\overset{+2}{\mathrm{C}}\,\overset{-2}{\mathrm{O}}\mathrm{(g)} \rightarrow \mathrm{2Fe(s)} + 3\,\overset{+4}{\mathrm{C}}\,\overset{-2}{\mathrm{O_2}}\mathrm{(g)}$$

$Fe^{+3} + 3e^- \rightarrow Fe$

Here, the O.N. of Fe decreases from +3 in $\mathrm{Fe_2O_3}$ to *zero* in Fe or addition of electron involved so $\mathrm{Fe_2O_3}$ is reduced to Fe.

$\overset{+2}{\mathrm{C}}\,\overset{-2}{\mathrm{O}}\mathrm{(g)} \rightarrow \overset{+4}{\mathrm{C}}\,\overset{-2}{\mathrm{O_2}}\mathrm{(g)}$the O.N. of C increases from +2 in CO to +4 in $\mathrm{CO_2}$ i.e., CO is oxidized to $\mathrm{CO_2}$.

In the above reaction oxidation and reduction, both is taking place simultaneously Hence, this reaction is a redox reaction.

(c) $4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3(s)$

The O.N. of each element in the given reaction can be represented as:

$$4\overset{+3}{B}\overset{-1}{Cl_3}(g) + 3\overset{+1}{Li}\overset{+3}{Al}\overset{-1}{H_4}(s) \rightarrow 2\overset{-3}{B_2}\overset{+1}{H_6}(g) + 3\overset{+1}{Li}\overset{-1}{Cl}(s) + 3\overset{+3}{Al}\overset{-1}{Cl_3}(s)$$

In this reaction, the O.N. of B decreases from $+3$ in BCl_3 to -3 in B_2H_6. i.e., BCl_3 is reduced toB_2H_6. Also, the O.N. of H increases from -1 in $LiAlH_4$ to $+1$ in B_2H_6 i.e., $LiAlH_4$ is oxidized to B_2H_6. In the above reaction oxidation and reduction, both is taking place simultaneously Hence, this reaction is a redox reaction.

(d) $2K(s) + F_2(g) \rightarrow 2K^+F^-(s)$

The O.N. of each element in the given reaction can be represented as:

$$2\overset{0}{K}(s) + \overset{0}{F_2}(g) \rightarrow 2K^+\overset{-1}{F}{}^-(s)$$

In this reaction, the O.N. of K increases from $zero$ in K to $+1$ in KF. i.e., K is oxidised to KF. Also, the O.N. of F_2 decreases from 0 in F_2 $and -1$ in KF i.e.redution is occure .

In the above reaction oxidation and reduction, both is taking place simultaneously Hence, this reaction is a redox reaction.

(e) $4\ NH_3 + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

The O.N. of each element in the given reaction can be represented as:

4

$$4\overset{-3}{N}\overset{+1}{H_3} + 5\overset{0}{O_2}(g) \rightarrow 4\overset{+2}{N}\overset{-2}{O}(g) + 6\overset{+1}{H_2}\overset{-2}{O}(g)$$

Here, the O.N. of N increases from -3 in NH_3 to $+2$ in NO. so oxidation is taking place there

the O.N. of O_2 decreases from 0 in O_2 to -2 in NO and H_2O i.e., so the gain of electron is occur here so O_2 is reduced.

In the above reaction oxidation and reduction, both is taking place simultaneously. Hence, this 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.

Solution:

Let us write the O.N. of each atom involved in the given reaction above its symbol as:

$$\overset{+1}{H}_2\overset{-2}{O} + \overset{0}{F}_2 \rightarrow \overset{+1}{H}\overset{-1}{F} + \overset{+1}{H}\overset{-2}{O}\overset{+1}{F}$$

Here, we have observed that the O.N. of F increases from 0 in F_2 to $+1$ in HOF.

Also, the O.N. of fluorine decreases from 0 in F_2 to -1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

5. Calculate the O.N. 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.

Solution:

(i) average O.N. of sulphure in H_2SO_5

Let the O.N. of S = x.

We know that,

The O.N. of $O = -2$

The O.N. of $H = +1$

Now,

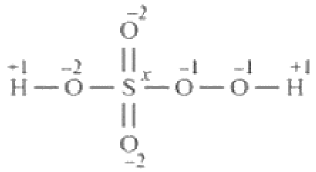
$$2(+1) + 1(x) + 5(-2) = 0$$

$$\Rightarrow 2 + x - 10 = 0$$

$$\Rightarrow x = +8$$

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

The structure of H_2SO_5 is shown as follows:



Now, $2(+1) + 1(\mathrm{x}) + 3(-2) + 2(-1) = 0$

$\Rightarrow\ 2 + \mathrm{x} - 6 - 2 = 0$

$\Rightarrow \mathrm{x} = +6$

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

(ii) $Cr_2O_7^{2-}$

Let the O.N. of Cr = x.

We know that,

The O.N. of $O = -2$

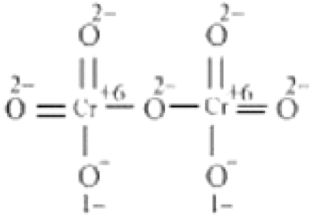
$2(\mathrm{x}) + 7(-2) = -2$

$\Rightarrow 2\mathrm{x} - 14 = -2$

$\Rightarrow \mathrm{x} = +6$

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

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



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

(iii) NO_3^-.

Let the O.N. of N = x.

We know that,

The O.N. of $O = -2$

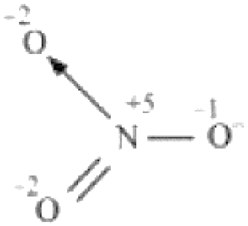
$1(\mathrm{x}) + 3(-2) = -1$

$\Rightarrow \mathrm{x} - 6 = -1$

$\Rightarrow x = +5$

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

The structure of NO_3^- is shown as follows:



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

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

Solution:

(a) Mercury(II) chloride: Mercury(II)=Hg^{+2} and chloride = Cl^- so the formula of a compound is $HgCl_2$

(b) Nickel(II) sulphate: nickel(II)=Ni^{+2} and sulphate =SO_4^{-2} so the formula of a compound is $NiSO_4$

(c) Tin(IV) oxide: $Tin(IV) = Sn^{+4} +$ and oxide =O^{-2} so the formula of a compound is SnO_2

(d) Thallium(I) sulphate:

$\text{Thallium(I)} = \text{Tl}^{+1}$

sulphate = SO_4^{-2}

the formula of a compound is TI_2SO_4

(e) Iron(III) sulphate: Iron(III) $= Fe^{+3}$ + and sulphate $= SO_4^{-2}$ so the formula of a compound is $Fe_2(SO_4)_3$

(f) Chromium(III) oxide: Chromium(III) $= Cr^{+3}$ + and oxide $= O^{-2}$ so the formula of a compound is 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.

Solution:

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

Substance	**O.N. of carbon**
CF_2CF_2, CCl_2CCl_2.COF_2, $COCl_2$, $COBr_2$	0
$FC \equiv CF$, $ClC \equiv CCl$	+1
$HC \equiv CH$	−1
$CHCl_3$, CHF_3, CO, (CF_2 CCl_2) carbeen	+2
CH_3F, CH_3Cl	−2
$Cl_3C - CCl_3$	+3
$H_3C - CH_3$	−3
CF_4, 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 carbon**
N_2	0
N_2O, $H_2N_2O_2$	+1
N_2H_2	−1
NO	+2
N_2H_4	−2
N_2O_3, HNO_2, NCl_3	+3

NH_3	-3
NO_2, N_2O_4	$+4$
N_2O_5, HNO_3	$+5$

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

Solution:

In sulphur dioxide (SO_2), the O.N. (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.

$SO_2 \rightarrow SO_3$(oxidation)

$SO_2 \rightarrow S^{-2}$(reduction)

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 O.N.s $+1$ and $+2$. Hence, H_2O_2 can act as an oxidising as well as a reducing agent.

$H_2O_2 \rightarrow O_2$(oxidation)

$H_2O_2 \rightarrow H_2O$ (reduction)

In ozone , the O. N. of Oxygen is zero and the range of the O. N. that O can have is from $+2$ to -2.

Therefore, the O. N. of O can only decrease in this case. Hence, O_3 acts as an oxidant.

$$O_3 \rightarrow O_2 + [O]$$

Ozone reduces peroxides to oxides and in turn gets reduced to oxygen so used as reductant .

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.

9. Consider the reactions:

(a) $6CO_2(g) + 6H_2O(l) \rightarrow C_6\,H_{12}O_6(aq) + 6O_2(g)$

(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.

Solution:

(a) The process of photosynthesis involves two steps.

Reaction 1

H_2O decomposes to give H_2 and O_2.

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

Reaction 2:

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

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

As reaction 1× *by* 6 and add with reaction 2 the net reaction of the process is given as:

$2H_2O(l) \longrightarrow 2H_2\ (g) + O_2\ (g)\ ] \times\ 6$

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

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

It is more appropriate to write the reaction as given above because water molecules are also

produced in the process of photosynthesis.

The path of this reaction can be investigated by using radioactive H_2O_{18} in place of H_2O.

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

The given reaction involves two steps. First, O_3 decomposes to form O_2 and O. $O_3(g) \rightarrow O(g) + 2O_2(g)$

Second ,H_2O_2 *decomposees to form* O_2 and H_2O

$$H_2O_2(l) + O(g) \rightarrow H_2O(l) + O_2(g)$$

Why it is more appropriate to write these reactions as :

In the second step, H_2O_2 reacts with the O produced in the first step, thereby producing H_2O and O_2.by adding both above 1and 2 reactions.

$$O_3(g) \rightarrow O(g) + 2O_2(g)$$

$$H_2O_2(l) + O(g) \rightarrow H_2O(l) + O_2(g)$$

Net reaction$H_2O_2(l) + O_3(g) \rightarrow H_2O(l) + O_2(g) + O_2(g)$

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

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

Solution:

The oxidation state of Ag in AgF_2 is $+2$. But, $+2$ is an unstable oxidation state of Ag. $+1$ is more stable oxidation state of Ag

Therefore, whenever AgF_2 is formed, silver readily accepts an electron to form Ag^+.

$$Ag^{+2} + e^- \rightarrow Ag^+$$

This helps to bring the oxidation state of Ag down from $+2$ to a more stable state of $+1$. Ag^{+2} reduced in Ag^+so 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.

Solution:

Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. This can be illustrated as follows:

(i) P_4 and F_2 are reducing and oxidising agents respectively.

If an excess of P_4 is treated with F_2 , then PF_3 will be produced, where the O.N. of P is +3 in PF_3

$$P_4(\text{excess}) + F_2 \longrightarrow PF_3$$

$O.No.of\ P$ 0 0 +3

However, if P_4 is treated with an excess of F_2, then PF_5 will be produced, wherein the O. N. of P is +5 in PF_5.

$$P_4(\text{excess}) + F_2 \longrightarrow PF_5$$

$O.No.of\ P$ 0 0 +5

(ii) K acts as a reducing agent, whereas O_2 is an oxidising agent. If an excess of K reacts with O_2, then K_2O will be formed, wherein the O. N. of O is –2 in K_2O.

$$4K(\text{excess}) + O_2 \longrightarrow 2K_2O$$

$O.No\ of\ O$ 0 0 -2

However, if K reacts with an excess of O_2, then K_2O_2 will be formed, wherein the O. N. of O is –1 in K_2O_2.

$$2K + O_2(\text{excess}) \longrightarrow K_2O_2$$

$O.No.of\ O$ 0 0 -1

(iii) C is a reducing agent, while O_2 acts as an oxidising agent.

If an excess of C is burnt in the presence of limiting amount of O_2, then CO will be produced, where the O. N. of Cabon is +2.

$$C(\text{excess}) + \frac{1}{2}O_2 \longrightarrow CO$$

$O.No.of\ C$ 0 0 +2

On the other hand, if C is burnt in an excess of O_2, then CO_2 will be produced, wherein the O. N. of C is +4.

$$C + O_2(\text{excess}) \longrightarrow CO_2$$

$O.No.of\,C$ 0 0 +4

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? \

Solution:

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

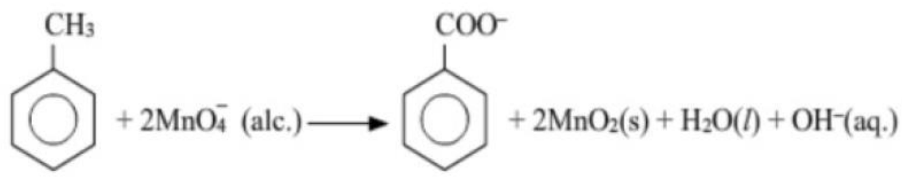
(i) In a neutral medium, OH^- ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

$$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4\,OH^-$$

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

Hence, in alcohol, $KMnO_4$ and toluene can react at a faster rate.

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



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

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

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

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

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

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

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

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

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

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

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

Solution:

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

Reduction : $e^- + Ag^+ \rightarrow Ag(s)$

Oxidation : $C_6H_6O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H^+ + 2e^-$

O.No of C 1/3 0

Oxidised substance → $C_6H_6O_2$

Reduced substance → AgBr

Oxidising agent → AgBr

Reducing agent → $C_6H_6O_2$

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

Oxidation : $H_2O(l) + HCHO(l) \rightarrow HCOO^-(aq) + 3H^+ + 2e^-$

Reduction : $[\mathrm{Ag\,(NH_3)_2}]^+(\mathrm{aq}) + e^- \rightarrow \mathrm{Ag(s)} + 2\mathrm{NH_3(aq)} + 2\mathrm{H_2O(l)}$

Oxidised substance → HCHO

Reduced substance → $[\mathrm{Ag\,(NH_3)_2}]^+$

Oxidising agent → $[\mathrm{Ag\,(NH_3)_2}]^+$

Reducing agent → HCHO

(c) $\mathrm{HCHO(l)} + 2\mathrm{Cu^{2+}\,(aq)} + 5\,\mathrm{OH^-(aq)} \rightarrow \mathrm{Cu_2O(s)} +$
$\mathrm{HCOO^-(aq)} + 3\mathrm{H_2O(l)}$
Reduction : $e^- + \mathrm{Cu^{2+}\,(aq)} \rightarrow \mathrm{Cu^{1+}\,(aq)}$
1. Oxidation: $\mathrm{HCHO(l)} \rightarrow \mathrm{HCOO^-(aq)}$
O.no. of carbon 0 +2

Oxidised substance → HCHO

Reduced substance → $\mathrm{Cu^{2+}}$

Oxidising agent → $\mathrm{Cu^{2+}}$

Reducing agent → HCHO

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

Oxidation : $\mathrm{N_2H_4(l)} \rightarrow \mathrm{N_2(g)}$

Oxidation No.of nitrogen -2 0

Reduction : $\mathrm{H_2O_2(l)} \rightarrow \mathrm{H_2O(l)}$

O. No.of oxygen -1 -2

Oxidised substance → $\mathrm{N_2H_4}$

Reduced substance → $\mathrm{H_2O_2}$

Oxidising agent → $\mathrm{H_2O_2}$

Reducing agent → $\mathrm{N_2H_4}$

(e) $\mathrm{Pb(s)} + \mathrm{PbO_2(s)} + 2\mathrm{H_2SO_4(aq)} \rightarrow 2\mathrm{PbSO_4(s)} + 2\mathrm{H_2O(l)}$
Oxidation : $\mathrm{Pb(s)} \rightarrow \mathrm{Pb^{+2}(s)} + 2\mathrm{e^-}$
Reduction : $2\mathrm{e^-} + \mathrm{Pb^{+4}(s)} \rightarrow \mathrm{Pb^{+2}(s)}$

Oxidised substance → Pb

Reduced substance → $\mathrm{PbO_2}$

Oxidising agent → $\mathrm{PbO_2}$

Reducing agent → Pb

14. Consider the reactions :

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

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

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

Solution:

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

Oxidation : $2S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2e^-$

$S_2O_3^{2-}$ thiosulphate get oxidized so it is reducing agent

Reduction : $2e^- + I_2(s) \rightarrow 2I^-(aq)$

I_2 get reduced so it is oxidising agent

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

oxidation : $S_2O_3^{2-}(aq) \rightarrow 2SO_4^{2-}(aq)$

O.No. of Sulphur +2 +6

reduction : $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$

$S_2O_3^{2-}$ get oxidized so it is reducing agent while Br_2 get reduced to it oxidizing agent here.

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

Solution:

F_2 can oxidize Cl^- to Cl_2, Br^- to Br_2 , and I^- to I_2

$F_2 + 2Cl^- \rightarrow 2\,F^{-1} + Cl_2$,

$F_2 + 2Br^- \rightarrow Br_2 + 2\,F^{-1}$

$F_2 + 2I^- \rightarrow I_2 + 2\,F^{-1}$

On the other hand, Cl_2, Br_2, and I_2 cannot oxidize F^- to F_2. The oxidizing power of halogens

increases in the order of $I_2 < Br_2 < Cl_2 < F_2$. Hence, fluorine is the best oxidant among halogens. Standard Reduction Potential value is highest for F_2

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

$2HI + H_2SO_4 \longrightarrow I_2 + SO_2 + 2H_2O$

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

I^- can reduce Cu^{2+} to Cu^+

$4I^-(aq) + 2Cu^{2+}(aq) \longrightarrow Cu_2I_2(s) + I_2(aq)$

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

Thus, the reducing power of hydrohalic acids increases in the order of $HF < HCl < HBr < HI$. Due to top to bottom reducing character increases as to bond length increase.

16. Why does the following reaction occur?

$XeO_6^{4-}(aq) + 2F^-(aq) + 6H^+(aq) \rightarrow XeO_3(g) + F_2(g) + 3H_2O(l)$

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

Solution:

The given reaction occurs because XeO_6^{4-} oxidises F^- and F^- reduces XeO_6^{4-}

$XeO_6^{4-}(aq) + 2F^-(aq) + 6H^+(aq) \rightarrow XeO_3(g) + F_2(g) + 3H_2O(l)$

In this reaction, the O.N. (O.N.) of Xe decreases from +8 in XeO_6^{4-} to +6 in XeO_3 and the O.N. of F increases from – 1 in F^- to *zero* in F_2 Hence, we can

conclude that Na_4XeO_6 is a stronger oxidising agent than F^- .

17. Consider the reactions:

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

(B) $H_3PO_2(aq) + 2CuSO_4(aq) + 2H_2O(l) \rightarrow H_3PO_4(aq) + 2Cu(s) + H_2SO_4(aq)$

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

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

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

Solution:

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

(B) $Cu^{+2}(aq) + 2e^- \rightarrow Cu(s)$

(C) $[Ag(NH_3)_2]^+(aq) + e^- \rightarrow Ag(s)$

Ag^+ and Cu^{2+} act as oxidising agents in reactions (a) and (b)respectively.

In reaction (c), Ag^+ oxidises C_6H_5CHO to $C_6H_5COO^-$, but in reaction (d), Cu^{2+} cannot oxidise C_6H_5CHO.

Hence, Cu^{2+} is a stronger oxidising agent than Cu^{2+}.

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

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

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

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

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

Solution:

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

Oxidation half reaction: $\mathrm{I^{-1}(aq) \rightarrow I_2(s)}$

Reduction half reaction: $Mn\mathrm{O_4^-(aq) \rightarrow M\ nO_2(aq)}$

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

Now, to balance the atom, we multiply $\mathrm{I^-}$ $with$ 2 into the LHS of the reaction

$\mathrm{2I^-(aq) \rightarrow I_2(s)}$

Now, to balance the charge, we add $\mathrm{2e^-}$ to the RHS of the reaction.

$\mathrm{2I^-(aq) \rightarrow I_2(s) + 2e^-}$

Step 3: In the reduction half-reaction, Balance the atoms other than O and

H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+

to balance H atoms.

Thus,we get : $\mathrm{MnO_4^-(aq) + 4H^+ \rightarrow MnO_2(aq)} + 2H_2O$

Step 4: The reaction occurs in a basic medium, so we add $\mathrm{4OH^-}$ ions to the RHS and right-hand side.

$\mathrm{MnO_4^-(aq) + 4H^+ + 4OH^- \rightarrow MnO_2(aq)} + 2H_2O\mathrm{+4OH^-}$

= $\mathrm{MnO_4^-(aq)} + 2H_2O \rightarrow \mathrm{MnO_2(aq) + 4OH^-}$

Step 5 : To balance the charge add 3 electrons are added to the LHS of the reaction.

$\mathrm{3e^- + MnO_4^-(aq)} + 2H_2O \rightarrow \mathrm{MnO_2(aq) + 4OH^-}$

Step 6: Equalising the number of electrons by multiplying the oxidation half-reaction by 3 and the reduction half-reaction by 2, we have:

$\mathrm{6I^-(aq) \rightarrow 3I_2(s) + 6e^-}$

$\mathrm{2MnO_4^-(cq) + 4H_2O + 6e^- \rightarrow MnO_2(s) + 8\ OH^-(aq)}$

Step 7: Adding the two half-reactions, we have the net balanced redox reaction as:

$$\mathrm{6I^-(aq) + 2MnO_4^-(aq) + 4H_2O(l) \rightarrow 3I_2(s) + 2MnO_2(s) + 8OH^-(aq)}$$

(b) $\mathrm{MnO_4^-(aq) + SO_2(g) \rightarrow Mn^{2+}(aq) + HSO_4^-(aq)}$(in acidic solution)Following the steps as in part (a), we have the oxidation half-reaction as:

$\mathrm{SO_2(g) \rightarrow HSO_4^-(aq)}$

Step 1: In the reduction half-reaction, Balance the atoms other than O and

H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+ to balance H atoms.

$2H_2O(l) + SO_2(g) \rightarrow HSO_4^-(aq) + 3H^+$

Step 2: To balance the charge add 2 electrons are added to the RHS of the reaction.

$2H_2O(l) + SO_2(g) \rightarrow HSO_4^-(aq) + 3H^+ + 2e^-$

the reduction half reaction as:

Step 3: In the reduction half-reaction, Balance the atoms other than O and

H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+ to balance H atoms.

$MnO_4^-(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Step 4 : To balance the charge add 5 electrons are added to the LHS of the reaction.

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

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

$$2MnO_4^-(aq) + 5SO_2(g) + 2H_2O(l) + H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5HSO_4^-(aq)$$

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

Step 1: To balance the charge add 1 electron in RHS

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

And the reduction half reaction as:

Step 2: Balance the atoms other than O and H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+ to balance H atoms.

$H_2O_2(aq) + 2H^+(aq) \rightarrow 2H_2O(l)$

Step 3: To balance the charge add 2 electrons are added to the LHS of the reaction.

$$2e^- + H_2O_2(aq) + 2H^+(aq) \rightarrow 2H_2O(l)$$

Step 4: Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:

$$H_2O_2(aq) + 2Fe^{2+}(aq) + 2H^+(aq) \rightarrow 2Fe^{3+}(aq) + 2H_2O(l)$$

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

Step 1: Balance the atoms other than O and H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+ to balance H atoms.

$$SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-} + 4H^+(aq)$$

Step 2: To balance the charge add 2 electrons are added to the RHS of the reaction.

$$SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-} + 4H^+(aq) + 2e^-$$

the reduction half reaction as:

Step 3: Balance the Cr atom multiply the $Cr^{3+}$with 2 in RHS

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}(aq)$$

Step 4: Balance the atoms other than O and H in each half-reaction individually. Here the to balance the atoms add H_2O to balance O atoms and H^+ to balance H atoms.

$$Cr_2O_7^{2-} + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

Step 5: To balance the charge add 6 electrons are added to the LHS of the reaction.

$$Cr_2O_7^{2-} + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

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

$$Cr_2O_7^{2-} + 3SO_2(g) + 2H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-} + H_2O(l)$$

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

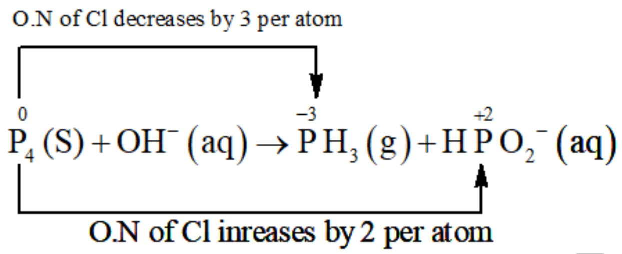
$P_4(s) + OH^-(aq) \rightarrow PH_3(g) + HPO_2^-(aq)$

$N_2H_4(l) + ClO_3^-(aq) \rightarrow NO(G) + Cl^-(G)$

$Cl_2O_7(g) + H_2O_2(aq) \rightarrow ClO_2^-(aq) + O_2(g) + H^+$

Solution:

(a)



The O.N. (O.N.) of P decreases from 0 in P_4 to – 3 in PH_3 and increases from 0 in P_4 to +2 in HPO_2^-. Hence, P_4 acts both as an oxidizing agent and a reducing agent in this reaction.

Ion–electron method:

The oxidation half equation is:

$P_4(s) \rightarrow HPO_2^-(aq)$

The P atom is balanced as:

$P_4 \rightarrow 4HPO_2^-(aq)$

The O atom is balanced by adding $8H_2O$ molecules:

$P_4 + 8H_2O \rightarrow 4HPO_2^-(aq)$

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

$P_4 + 8H_2O \rightarrow 4HPO_2^-(aq) + 12H^+$

The charge is balanced by adding 8 e^- as:

$P_4 + 8H_2O \rightarrow 4HPO_2^-(aq) + 12H^+ + 8e^- \ldots(i)$

The reduction half equation is:

$P_4(s) \rightarrow PH_3(g)$

The P atom is balanced as:

$P_4 \rightarrow 4PH_3(g)$

The H is balanced by adding $12H^+$ as:

$P_4 + 12H^+ \rightarrow 4PH_3(g)$

The charge is balanced by adding $12e^-$ as:

$P_4 + 12H^+ + 12e^- \rightarrow 4PH_3(g)$…(ii)

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

$$5P_4(s) + 24H_2O \rightarrow 12HPO_2^- + 8PH_3(g) + 12H^+$$

As, the medium is basic, add $12OH^-$ both sides as:

$5P_4(s) + 12H_2O + 12OH^- \rightarrow 12HPO_2^- + 8PH_3(g)$

This is the balanced equation by ion-electron method

O.N. method:

Let, total no of P reduced $=\ x$

$\therefore$Total no of P oxidised $=\ 4 - x$

$P_4(s) + OH^- \rightarrow xPH_3(g) + (4 - x)HPO_2^-$…(i)

Total decrease in O.N. of P $= x \times 3 = 3x$

Total increase in O.N. of P $= (4 - x) \times 2 = 8 – 2x$

$\because 3x = 82x\ x = 8/5$ from (i),

$5P_4(s) + 5OH^- \rightarrow 8PH_3(g) + 12HPO_2^-$

Since, reaction occurs in basic medium, the charge is balanced by adding $7OH^-$ on LHS as:

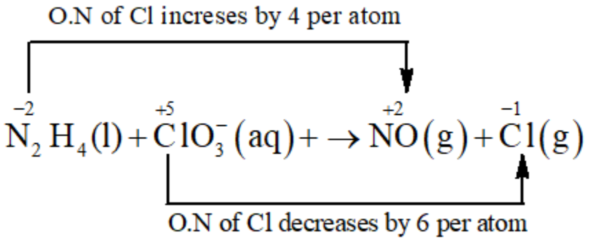
$5P_4(s) + 12OH^- \rightarrow 8PH_3(g) + 12HPO_2^-$

The O atoms are balanced by adding $12H_2O$ as:

$5P_4(s) + 12H_2O + 12OH^- \rightarrow +12HPO_2^- + 8PH_3(g)$

This is the required balanced equation.

(b)



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

Ion–electron method:

The oxidation half equation is:

$\overset{-2}{N_2}H_4(l) \rightarrow \overset{+2}{N}O(g)$

The N atoms are balanced as:

$N_2H_4(l) \rightarrow 2NO(g)$

The O.N. is balanced by adding 8 electrons as

$N_2H_4(l) \rightarrow 2NO(g) + 8e^-$

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

$N_2H_4(l) + 8OH^-(aq) \rightarrow 2NO(g) + 8e^-$

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

$N_2H_4(l) + 8OH^-(aq) \rightarrow 2NO(g) + 6H_2O(l) + 8e^- \ldots$(i)

The reduction half equation is:

$\overset{+2}{C}lO_2^-(aq) \rightarrow \overset{+1}{C}l^-(aq)$

The O.N. is balanced by adding 6 electrons as:

$ClO_3^-(aq) + 6e^- \rightarrow Cl^-(aq)$

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

$ClO_3^-(aq) + 6e^- \rightarrow Cl^-(aq) + 6OH^-(aq) \ldots$(ii)

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

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow 6NO(g) + 4Cl^-(aq) + 6H_2O(l)$

O.N. method:

Total decrease in O.N. of N $= 2 \times 4 = 8$

Total increase in O.N. of Cl $= 1 \times 6 = 6$

On multiplying N_2H_4 with 3 and ClO_3^- with 4 to balance the increase and decrease in O. N., we get:

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow NO(g) + Cl^-(aq)$

The N and Cl atoms are balanced as:

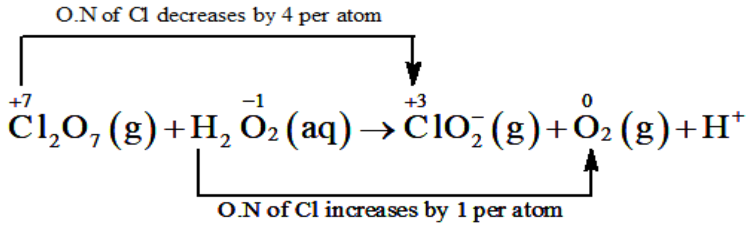
$3N_2H_4(l) + 4ClO_3(aq) \rightarrow 6NO(g) + 4Cl^-(aq)$

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

$3N_2H_4(l) + 4ClO_3^-(aq) \rightarrow 6NO(g) + 4Cl^-(aq) + 6H_2O(l)$

This is the required balanced equation.

(c)



The O.N. of Cl decreases from +7 in Cl_2O_7 to +3 in ClO_2^- and the oxidation

number of O increases from – 1 in H_2O_2 to zero in O_2. Hence, in this reaction, Cl_2O_7 is the

oxidizing agent and H_2O_2 is the reducing agent.

Ion–electron method:

The oxidation half equation is:

$H_2\overset{-1}{O}_2(aq) \rightarrow \overset{0}{O}_2(g)$

The O.N. is balanced by adding 2 electrons as:

$H_2O_2(aq) \rightarrow O_2(g) + 2e^-$

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

$H_2O_2(aq) \rightarrow 2OH^-(aq) + 2e^-$

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

$H_2O_2(aq) + 2OH^-(aq) \rightarrow +2H_2O(l) + 2e^-$ …(i)

The reduction half equation is:

$\overset{+7}{C}l_2O_7(g) \rightarrow \overset{+3}{C}lO_2^-(aq)$

The Cl atoms are balanced as:

$Cl_2O_7(g) \rightarrow 2ClO_2^-(aq)$

The O.N. is balanced by adding 8 electrons as:

$Cl_2O_7(g) + 8e^- \rightarrow 2ClO_2^-(aq)$

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

$Cl_2O_7(g) + 8e^- \rightarrow 2ClO_2^-(aq) + 6OH^-(aq)$

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

$Cl_2O_7(g) + 3H_2O(l)8e^- \rightarrow 2ClO_2^-(aq) + 6OH^-(aq)$ …(ii)

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

$Cl_2O_7(g) + 4H_2O_2(aq) + 2OH^-(aq) \rightarrow 2ClO_2^- + 4O_2(g) + 5H_2O(l)$

O.N. method:

Total decrease in O.N. of $Cl_2O_7 = 4 \times 2 = 8$

Total increase in O.N. of $H_2O_2 = 2 \times 1 = 2$

By multiplying H_2O_2 and O_2 with 4 to balance the increase and decrease in the O.N., we get:

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow ClO_2^-(aq) + 4O_2(g)$

The Cl atoms are balanced as:

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow 2ClO_2^-(aq) + 4O_2(g)$

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

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow 2ClO_2^-(aq) + 4O_2(g) + 3H_2O(l)$

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

$Cl_2O_7(g) + 4H_2O_2(aq) \rightarrow 2ClO_2^-(aq) + 4O_2(g) + 5H_2O(l)$

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

$$(CN)_2(g) + 2OH^-(aq) \rightarrow CN^-(aq) + CNO^-(aq) + H_2O(l)$$

Solution:

Let the O.N. of C be x.

$(CN)_2$

$2(x-3) = 0$

$\therefore x = 3$

CN^-

$x - 3 = -1$

$\therefore x = 2$

CNO^-

$x - 3 - 2 = -1$

$\therefore x = 4$

The O.N.s of carbon in $(CN)_2$, CN^- and CNO^- are $+3$, $+2$ and $+4$

The O.N. of carbon in the various species is:

$$\left(\overset{+3}{C}N\right)_2(g) + 2OH^-(aq) \rightarrow \overset{+2}{C}N^-(aq) + \overset{+4}{C}NO^-(aq) + H_2O(l)$$

It can be easily observed that the same compound $\left(\overset{+3}{C}N\right)_2$ is being reduced in $\overset{+2}{C}N^-$ and oxidized in $\overset{+4}{C}NO^-$ simultaneously in the given equation. So it is a disproportionation reaction. Thus, it can be said that the alkaline decomposition of cyanogen is an example of 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.

Solution:

The given reaction can be represented as:

$$Mn^{3+}(aq) \rightarrow Mn^{2+}(aq) + MnO_2(s) + H^+(aq)$$

The oxidation half equation is:

$Mn^{3+}(aq) \rightarrow MnO_2(s)$

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

$Mn^{3+} + 2H_2O(l) \rightarrow MnO_2(s) + 4H^+(aq)$

To balance the charge add 1 electron in RHS

$Mn^{3+} + 2H_2O(l) \rightarrow MnO_2(s) + 4H^+(aq) + e^-$ …(i)

The reduction half equation is:

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

To balance the charge add 1 electron in LHS

$Mn^{3+}(aq) + e^- \rightarrow Mn^{2+}(aq)$ …(ii)

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

$2Mn^{3+}(aq) + 2H_2O(l) \rightarrow MnO_2(s) + 2Mn^{2+}(aq) + 4H^+(aq)$

22. Consider the elements :

Cs, Ne, I and F

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

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

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

(d) Identify the element which exhibits neither the negative nor does the positive oxidation state.

Solution:

(a) F exhibits only negative oxidation state of -1 because it has highest electronegativity

(b) Cs exhibits positive oxidation state of $+1$because it is alkali metal atom

(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 can not exhibits negative or positive oxidation states because it has complete octate

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.

Solution:

The given redox reaction can be represented as:

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

The oxidation half-reaction is.

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

to balance the $O -$ atom add H_2O in oxygen deficient side

$2H_2O\ (l) + SO_2(aq) \rightarrow SO_4^{2-}(aq)$

to balance the $H -$ atom add H^+ in hydrogen deficient side

$2H_2O\ (l) + SO_2(aq) \rightarrow SO_4^{2-}(aq) + 4H^+(aq)$

to balance charge add add two electrons as:

$2H_2O\ (l) + SO_2(aq) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

$Cl_2(s) \rightarrow Cl^-(aq)$

The chlorine atoms are balanced as:

$Cl_2(s) \rightarrow 2Cl^-(aq)$

to balance charge add add two electrons as:

$2e^- + Cl_2(s) \rightarrow 2Cl^-(aq)$

Add both half reaction we get net reaction :

$Cl_2(s) + SO_2(aq) + 2H_2O(l) \rightarrow 2Cl^-(aq) + SO_4^{2-}(aq) + 4H^+(aq)$

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 can show disproportionation reaction.

Solution:

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

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

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

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?

Solution:

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

$$4\,NH_3 \quad + \quad 5\,O_2 \quad \rightarrow \quad 4NO \quad + \quad 6\,H_2O$$

Weight 10.00 g 20.00 g - -

Moles 10/17 20/32

$$\frac{Moles}{S.C.} \quad \frac{\frac{10}{17}}{4} = 0.147 \quad \frac{\frac{20}{32}}{5} = 0.104$$

O_2 is limiting reagent

Thus, 68 g of NH_3 reacts with 160 g of O_2

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

Therefore, O_2 is the limiting reagent (we have considered the amount of O_2 to calculate the weight of nitric oxide obtained in the reaction).

Now, 160g of O_2 gives 120g of NO.

Therefore, 20g of O_2 gives $\frac{120\times 20}{160}$ g of N, or 15g of NO.

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

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

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

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

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

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

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

Solution:

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

$$2Fe^{3+}(aq) + I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s)$$

Oxidation half equation: $2I^-(aq) \rightarrow I_2(s) + 2e^-$; $E^\circ = -0.54V$

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

$$\mathbf{2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s);\ E^\circ = +0.23V}$$

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

(b) The possible reaction between $Ag^+(aq) +\ Cu(s)$ is given by,

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

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

Reduction half equation: $\left[Ag^+(aq) + e^- \rightarrow Ag(s)\right]\times 2$; $E^\circ = +0.80V$

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

E^o for the overall reaction is positive. Thus, the reaction between $Ag^+(aq)$ and $Cu(s)$ is feasible.

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

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

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

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

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

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

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

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

Oxidation half equation: $Ag(s) \rightarrow Ag^+(aq) + e^-;$ $E^\circ = -0.80V$

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

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

Here E^o for the overall reaction is positive. Thus, the reaction between $Fe^{3+}(aq)$ and $Ag(s)$ is not feasible.

(e) The possible reaction between $Br_2(aq)$ and $Fe^{2+}(aq)$ is given by

$Br_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Br^-(aq) + 2Fe^{3+}(aq)$

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

Reduction half equation: $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq);$ $E^\circ = +1.09V$

$Br_2(aq) + 2Fe^{2+}(aq) \rightarrow 2Br^-(aq) + 2Fe^{3+}(aq);\ E^\circ = +0.23V$

Here E^o for the overall reaction is positive. Thus, 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.

Solution:

(i) $AgNO_3$ ionizes in aqueous solutions to form Ag^+ and NO_3^- ions. On electrolysis,either Ag^+ ions or H_2O molecules can be reduced at the cathode. But the reduction potential of Ag^+ ions is higher than that of H_2O.

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

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq); E^o = 0.83V$

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

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

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

Therefore, Ag metal gets oxidized at the anode.

(ii) Pt cannot be oxidized easily. Hence, at the anode, oxidation of water occurs to liberate O_2.

At the cathode, Ag + ions are reduced and get deposited.

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

$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}$

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

$2H^+(aq) + 2e^- \rightarrow H_2(g); E^o = 0.0V$

$2H_2O(aq) + 2e^- \rightarrow H_2(g)OH^-(aq); E^o = -0.83V$

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

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

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

$CuCl_2(aq) \rightarrow Cu^{2+}(aq) + 2Cl^-(aq)$

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

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

$H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-; E^o = -0.83V$

Hence, Cu^{2+} ions are reduced at the cathode and get deposited.

Similarly, at the anode, either of Cl^- or H2O is oxidized. The oxidation potential of H2O is

higher than that of Cl^-

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-; E^o = -1.36V$

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

But oxidation of H_2O molecules occurs at a lower electrode potential than that of Cl^- ions because of over-voltage (extra voltage required to liberate gas). As a result, Cl^- .ions are oxidized at the anode to liberate Cl_2 gas.

Similarly, at the anode, either of Cl^-or H_2O is oxidized. The oxidation potential of H_2O is higher than that of Cl^-.

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.

Solution:

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

According to electrochemical series the order of the increasing reducing the 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: $Mg > Al > Zn > Fe, > Cu$

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.

Solution:

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

30. Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$takes place, Further show:

(i) which of the electrode is negatively charged,

(ii) the carriers of the current in the cell, and

(iii) individual reaction at each electrode

Solution:

The galvanic cell corresponding to the given redox reaction can be represented 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) Ions are the carriers of current in the cell.

(iii) The reaction taking place at the Zn electrode can be represented as:

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

And the reaction taking place at the Ag electrode can be represented as:

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