CHAPTER: REDOX REACTION

Question 8.1 Assign oxidation numbers to the dark elements in each of the following species:
$NaH_2PO_4, NaHSO_4, K_2MnO_4, CaO_2, NaBH_4, H_2S_2O_7, KAl(SO_4)_2.12H_2O$

Answer : The oxidation number of an atom in an

molecular or in ion is defined as the number of charge

it would carry if the electrons were completely

transferred.(i)NaH_2PO_4

We know that sodium is a Earth metal so it's oxidation

number is $+1$

Oxidation number of hydrogen is $+1$

Oxidation number of oxygen is -2

Now the oxidation number of phosphorous is X

So now

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

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

$$X - 5 = 0$$

$$X = +5$$

Thus the oxidation number of phosphorus is +5 .

(ii)$NaHSO_4$

We know that Na Oxidation number$= +1$

Oxidation number of $H = +1$

Oxidation number of $O = -2$

Now the oxidation number of Sulphur is $= X$

So now ,

$$+1 + 1 + X - 8 = 0$$

$$2 + X - 8 = 0$$

$$X - 6 = 0$$

$$X = +6$$

Thus the oxidation number of Sulphur is $+6$

(iii) $H_4P_2O_7$

Oxidation number of $H = +1$

Oxidation number of $O = -2$

Lets assume that the oxidation number of $P = X$

So now ,

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

$$4 + 2X - 14 = 0$$

$$2X - 10 = 0$$

$$X = +5$$

Thus the oxidation number of $P = +5$

(iii)K_2MnO_4

Oxidation number of $k = +1$

Oxidation number of $O = -2$

Lets assume that the oxidation number of $Mn = X$

So now ,

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

$$2 + X - 8 = 0$$

$$X - 6 = 0$$

$$X = +6$$

Thus the oxidation number of $Mn = +6$

(v)CaO_2

Oxidation number of $Ca = +2$

Lets assume that the oxidation number of $O = X$

So now ,

$$(+2) + (X)2 = 0$$

$$2X = -2$$

$$X = -1$$

Thus the oxidation number of $O = -1$

(vi) $NaBH_4$

Oxidation number of $Na = +1$

Oxidation number of $H = -1$

Lets assume oxidation number of $B = X$

So now,

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

$$X - 3 = 0$$

$$X = +3$$

Thus the oxidation number of $B = +3$

(vii) $H_2S_2O_7$

Oxidation number of $\mathrm{H} = +1$

Oxidation number of $\mathrm{O} = -2$

Let's assume oxidation number of $\mathrm{S} = \mathrm{X}$

So now,

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

$$2 + 2\mathrm{X} - 14 = 0$$

$$2\mathrm{X} = +12$$

$$\mathrm{X} = +6$$

Thus the oxidation number of $\mathrm{S} = +$

$$\mathrm{KAl(SO_4)_2.12H_2O}$$

Oxidation number of $\mathrm{K} = +1$

Oxidation number of $\mathrm{Al} = +3$

Oxidation number of $\mathrm{O} = -2$

Oxidation number of $\mathrm{H} = +1$

Let's assume oxidation number of $\mathrm{S} = \mathrm{X}$

So now,

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

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

$$2\mathrm{X} = +12$$

$$\mathrm{X} = +6$$

Thus the oxidation number of $\mathrm{S} = +6$

Question 8.2 What are the oxidation numbers of the Dark elements in each of the following and How do you rationalize your results?

$$\mathrm{KI_3, H_2S_4O_6, Fe_3O_4, CH_3CH_2OH, CH_3COOH}$$

Answer: (i)$\mathrm{KI_3}$

Let's assume oxidation number of $\mathrm{I} = \mathrm{X}$

the oxidation number of $\mathrm{K} = +1$

$$1(+1) + 3(\mathrm{X}) = 0$$

$$+1 + 3\mathrm{X} = 0$$

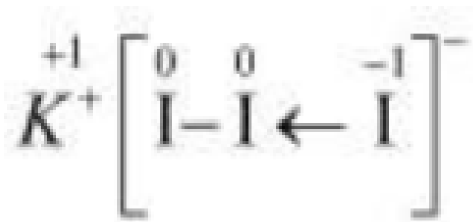
$$3\mathrm{X} = -1$$

$$\mathrm{X} = -1\,/\,3$$

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

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

So, in a this molecule, the oxidation number of the two I atoms forming the I^2 molecule is 0, whereas



the oxidation number of the I atom forming the coordinate bond is -1

(ii)H_2S_4O

Let assume oxidation number of S $= X$

The oxidation number of H $= +1$

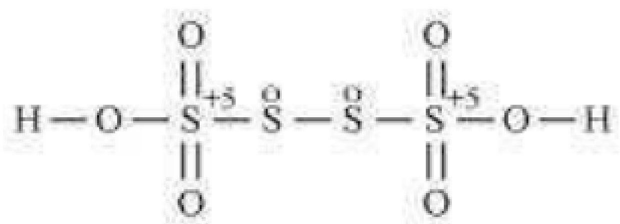
The oxidation number of O $= -1$

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

$$4X - 10 = 0$$

$$X = +10/4$$

SO oxidation number cannot be fractional. Hence, Smust be present in different oxidation states in the molecule.



So the oxidation number of two of the four S atoms is $+5$ and the oxidation number of the other two S atoms is 0.

(iii)Fe_3O_4

Let assume oxidation number of Fe $= \mathrm{X}$

The oxidation number of $\mathrm{O} = -2$

$$3(\mathrm{X}) + 4(-2) = 0$$

$$3\mathrm{X} - 8 = 0$$

$$\mathrm{X} = 8 / 3$$

Now ,oxidation number cannot be fractional.

So, one of the three Fe atoms exhibits the oxidation number of $+2$ and the other two Fe atoms exhibit the Oxidation number of $+3$.

$$\overset{+2}{\mathrm{Fe}}\mathrm{O}, \overset{+3}{\mathrm{Fe}}_2\mathrm{O}_3$$

(iv)$\mathrm{CH_3CH_2OH}$Le

Let assume oxidation number of $\mathrm{C} = \mathrm{X}$

The oxidation number of $\mathrm{O} = -2$

The oxidation number of $\mathrm{H} = +1$

$$\mathrm{X} + 3(+1) + \mathrm{X} + 2(+1) + 1(-2) + 1(+1) = 6$$

$$\mathrm{X} + 3 + \mathrm{X} + 2 - 2 + 1 = 0$$

$$2\mathrm{X} + 4 = 0$$

$$\mathrm{X} = -2$$

So the oxidation number of $\mathrm{C} = -2$

(iv)$\mathrm{CH_3COOH}$

Let assume oxidation number of $\mathrm{C} = \mathrm{X}$

The oxidation number of $\mathrm{O} = -2$

The oxidation number of H $= +1$

$$X + 3(+1) + X + (-2) + (-2) + 1(+1)$$

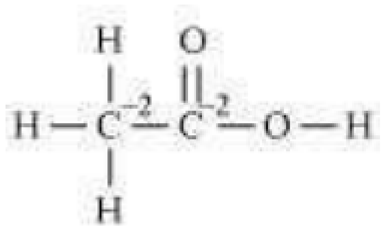
$$2X + 3 - 2 - 2 + 1 = 0$$

$$2X = 0$$

$$X = 0$$

So 0 is average oxidation number of C

The two carbon atoms present in this molecule are present in different environments. That's why, they cannot have the same oxidation number. Thus, C exhibits the oxidation states of $+2$and $+2$in CH_3COOH



Question8.3 Justify that the following reactions are redox reactions: (i)$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

$$\text{(ii)}Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3O_2(g)$$

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

$$\text{(iv)}2K(s) + F_2(g) \rightarrow 2K + \bar{F}(s)$$

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

Answer : (i)The given reaction is shown below.

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

In the given reaction, the oxidation number of copper is changing from $+2$ to $0$and that of hydrogen is changing from 0 to $+1$. It means copper is getting reduced while oxygen is getting oxidized. Thus, the given reaction is a redox reaction.

(ii)The given reaction is shown below.

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3O_2(g)$$

In the given reaction, the oxidation number of iron is changing from $+3$ to 0 and that of carbon is changing from $+2$ to -4. It means iron is getting reduced while carbon is getting oxidized. Thus, the given reaction is a redox reaction.

(iii)The given reaction is shown below.

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

In the given reaction, the oxidation number of boron is changing from $+3$ to -3 and that of hydrogen is changing -1 from $+1$ to . It means boron is getting reduced while hydrogen is getting oxidized. Thus, the given reaction is a redox reaction.

(iv)The given reaction is shown below.

$$2K(s) + F_2(g) \rightarrow 2K + \bar{F}(s)$$

In the given reaction, the oxidation number of potassium is changing from 0 to $+1$ and that of fluorine is changing from 0 to -1It means potassium is getting oxidised while fluorine is getting reduced. Thus, the given reaction is a redox reaction.

(v)The given reaction is shown below.

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

In the given reaction, the oxidation number of nitrogen is changing from -3 to$+2$ and that of oxygen is changing from 0 to -2. It means nitrogen is getting oxidised while oxygen is getting reduced. Thus, the given reaction is a redox reaction.

Question 8.4 Fluorine reacts with ice and results in the change:

$$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF$$

Justify that this reaction is a redox reaction.

Answer : First if all oxidation number of F in $F_2 = 0$

oxidation number of F in HF $= -1$

oxidation number of F in HOF $= +1$

$$\overset{+1}{H_2}\overset{-2}{O} + \overset{0}{F_2} \longrightarrow \overset{+1}{H}\overset{-1}{F} + \overset{+1}{H}\overset{-2}{O}\overset{+1}{F}$$

Here, we see oxidation number off decreases from 0 to -1(HF) and increase from $+1$(HOF). This shows that F_2 is both reduced as well as oxidised.

So , it is a redox reaction and more specifically , it is a dis proportionation redox reaction.

Question 8.5 Calculate the oxidation number of sulphur, chromium and nitrogen in –

$$H_2SO_5, Cr_2O^{2-}{}_7\,, NO_3{}^-$$

Suggest structure of these compounds. Count for the fallacy.

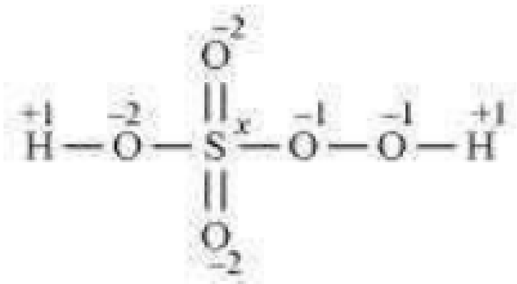
Answer : H_2SO_5 by conventional method.

Let assume X be the oxidation number of S.

$2(+1) + X + 5(-2) = 0 \; 2 + X - 10 = 0$

$$X = +8$$

Now S can't be 8 or more than 8 because S has only 6 valance electrons. This fallacy is removed by calculating oxidation of S by chemical bonding method.



Now H_2SO_5 two oxygen atoms are in -1 oxidation state.

Let see the oxidation number of $S = X$

$$2(+1) + X + 3(-2) + 2(-1) = 0$$

$$+2 + X - 6 - 2 = 0$$

$$X = +6$$

Here ,oxidation number of $S = +6$

$$\text{(ii)}Cr_2O_7{}^{2-}$$

Let assume the oxidation number of $Cr = X$

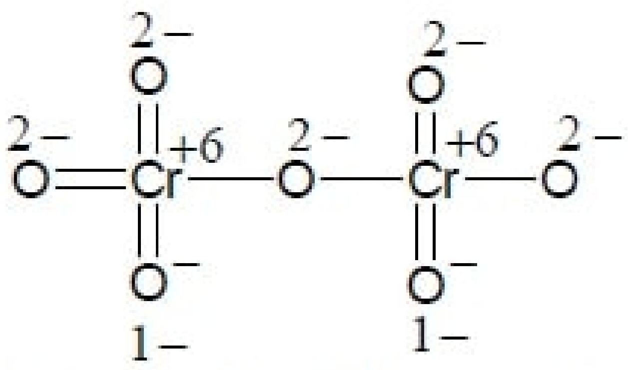
Now,

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

$$2X = 12$$

$$X = +6$$

Here , the oxidation number of Cr in dichromate ion is $+6$



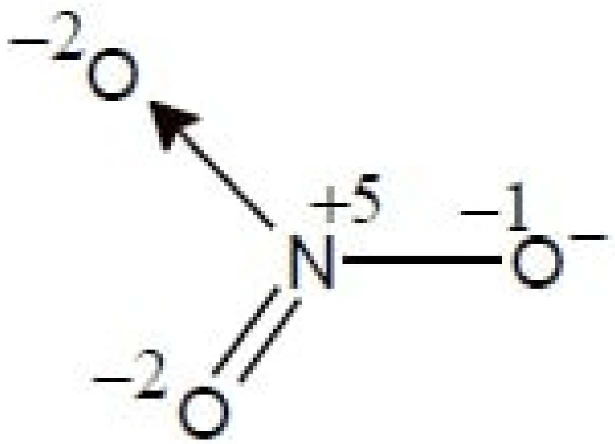
(iii)NO_3^-

Let assume X be the oxidation number of N in nitrate ion.

$$X + 3(-2) = -1$$

From the structure

$$X + 1(-1) + 1(-2) + 1(-2) = 0$$



$$X - 1 - 2 - 2 = 0$$

$$X = +5$$

Question 8.6 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(iv)oxide

Answer: (a) The formula of mercury (ii)chloride → $HgCl_2$

(b)The formula of Nickel (ii)sulphate → $NiSO_4$

(c) The formula of Tin(iv)oxide → SnO_2

(d) The formula of Thallium(i)sulphate → Ti_2SO_4

(e) The formula of Iron (iii)sulphate → $Fe_2(SO_4)_3$

(f) The formula of Chromium(iv)oxide $\rightarrow \mathrm{Cr_2O_3}$

Question 8.7 Suggest a list of the substances where carbon can exhibit oxidation states from -4 to $+4$ and nitrogen from -3 to $+5$

Answer: The state of an atom having a particular oxidation number.

OR The combination of a substance with oxygen .

The carbon compounds whose oxidation states are -4 to $+4$ are given in the following table.

Substance	O.N. of carbon
CH_2Cl_2	0
$ClC\equiv CCl$	+1
$HC\equiv CH$	-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

Nitrogen compounds whose oxidation states are -3 to $+5$ given 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 8.8 While Sulphur dioxide and hydrogen peroxide can act as oxidizing as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants Why?

Answer: Oxidation number of S in Sulphur dioxide –

Suppose , oxidation number of Sulphur = X

Oxidation number of oxygen =−2

So now ,

$$X + (-2)2 = 0$$

$$X = +4$$

So the oxidation state of Sulphur are +4 and − 1 Respectively , as these can increase as well as decrease when the compounds take part in chemical reactions ,hence they can act as oxidizing as well as reducing agents. Suppose the oxidation number of Oxygen in hydrogen peroxide = X Oxidation number of Hydrogen = +1

So now ,

$$2 + 2X = 0$$

$$X = -1$$

In hydrogen peroxide ,the oxidation number of O = −1and the range of the oxidation number that O Can have are from 0 to − 2 .Can sometimes also attain the oxidation number +1 and + 2 Hence, hydrogen peroxide can act as an oxidizing as well as reducing agents. We know that in ozone ,the oxidation state is 0, while in nitric acid the oxidation state of nitrogen is +5.As both can underdo decrease in oxidation state and not an increase in its value hence they can act only as oxidants and not as reductions .

Question 8.9 Consider the reactions:

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

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

Why it is more appropriate to write these reactions as:

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

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

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

Answer: First of all we have A reaction which is a synthesis reaction . and this reaction is completed in two steps.

First step – $H_2O\ (l) \rightarrow H_2(g) + 1/2\, O_2(g)$

This H_2 produced will react with carbon dioxide in step two and rede it into glucose.

Second step : $6CO_2 + 12H_2 \rightarrow C_6H_{12}O_6 + 6H_2O$

In both these steps, all reactions are processed, now we have to club both the reactions, for this, glucose, water, oxygen should come in the product side. For this, multiply with 12 to remove the extra product hydrogen , so now

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

$$\left.2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}\right] \times 6$$

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

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

produced in the process of photosynthesis . The path of this mechanism can be investigated by using radio action H_2O in place of O.

The given reaction involves two steps :

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

These oxygen steps heat up with hydrogen peroxide in seconds .

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

After this ,

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

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

Question 8.10 The compound AgF_2 is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

Answer: $AgF_2$is a stable compound, if it forms

compound, it will act as a strong oxidizing agent. So we know that Ag^2+ acts as $+2$ oxidation state so the electronic configuration of Ag will be something like this

$$Ag \rightarrow 4d^{10}, 5s^1$$

It will lose only one electron so the electronic

configuration of Ag will go to $4d^{10}$ which is stable. It will lose one electron to get Ag^2+ so the configuration will be $4d^9$ which are stable.Every species wants to be in a stable state whether it loses electrons or gains So it will have to show reduction reaction so that it will act as an oxidizing agent.

Question 8.11 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 agent is in excess and a compound of higher oxidation state is formed if the oxidizing 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 agent is in excess and a compound of higher oxidation state is formed if the oxidation state is in excess.Following three illustration justify

this –

(i)when P_4(excess) reacts with F_2after that PF_3 is produced in

which p has $3+$ oxidation number ,

Now oxidizing agent is F_2 and the reducing agent is P_4.

So,

$$P^4 + F^{2(\text{excess})} \rightarrow P^5 + F_5$$

When fluorine is in excess , PF_5 is formed in which P has an

oxidizing number of $+5$

$$P_4(\text{acess}) + F_2 \rightarrow P^3 + F_3$$

(ii)when (K excess) reacts with Oxygen after that K_2O is produced

in which oxygen has -2 oxidation number ,

Now oxidizing agent is oxygen and the reducing agent is K

So,

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

When Oxygen is in excess , K_2O_2 is formed in which Oxygen has

an oxidation number of -1

$$2K + O_2(\text{excess}) \rightarrow K_2O_2{}^1 -$$

(iii)when C(excess) reacts with Oxygen after that CO_2 is produced

in which carbon has $+4$ oxidation number ,

So,

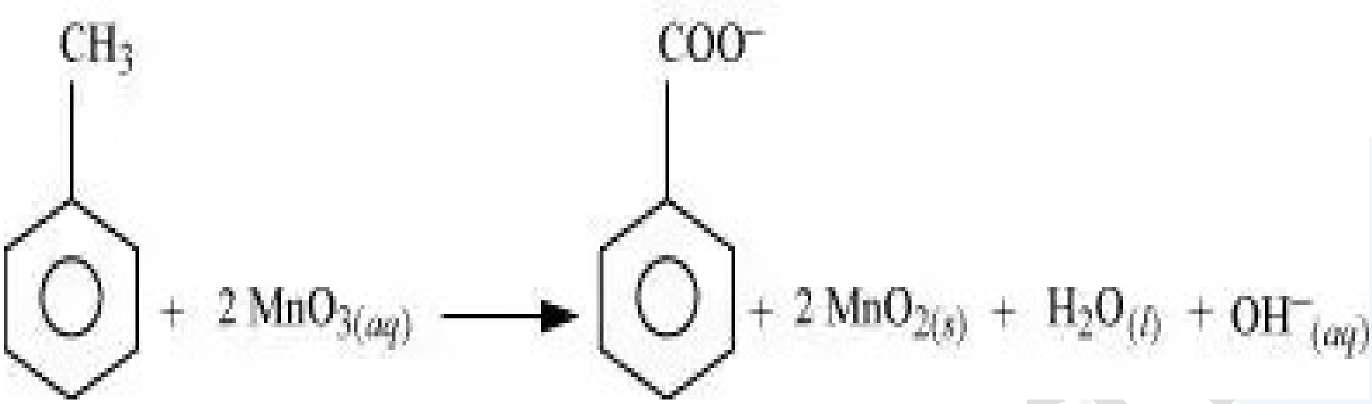
$$C + O_2(\text{excess}) \rightarrow CO_2 +$$

Question 8.12 How do you count for the following observations?

(a) 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 colorless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why?

Answer: (a) the reaction will ne produced in neutral medium and we know that in neutral medium OH^- ions are produced in reaction its self.As a result of this corresponding acid and base are reduced .So we have to add alcohol $KMnO_4$ instead of acidic $KMnO_4$



(b)We know that by adding solutions of bromide ions in NaBr we will get full which are strong reducing acids. Because of this he will reduce H_2SO_4 to H_2S.When we reduce H_2SO_4 solution in a compound which has chloride ions and mix it with H_2SO_4 then the following reaction will be

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

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

This HCl they are a reducing agent so it will not reduce H_2SO_4

so that we do not get any gas

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

Question 8.13 Identify the substance oxidized, reduced, oxidizing agent and reducing agent for each of the following reactions.

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

(b)$HCHO(l) + 2[Ag(NH_3)_2] + 3O\bar{H}(aq) \rightarrow 2Ag(s) + HCO\bar{O}(aq) + NH^3(aq) + 2H^2O$

(c)$HCHO(l) + 2Cu_2 + (aq) + 5O\bar{H}(aq) \rightarrow Cu_2O(s) + HCO\bar{O}(aq) +$ $3H_2O(l)$

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

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

Answer: Oxidizing agents means oxidation by accepting electrons and a reducing agents is a substance that causes reduction by loosing electrons .

(1) Oxidized substance→ $C_6H_6O_2$

Reduced substance → $AgBr$

Oxidizing agent → $AgBr$

Reducing agent → $C_6H_6O_2$

(2) in second reaction oxidized , reduced, oxidizing reagent , reducing agent are follows.

Oxidized substance → $HCHO$

Reduced substance → $[Ag(NH_3)_2]$

Oxidizing agent → $[Ag(NH_3)_2]$

Reducing agent → $HCHO$

(3) in third reaction oxidized, reduced, oxidizing agent , reducing agent are follows.

Oxidized substance → $HCHO$

Reduced substance → $Cu^2 +$

Oxidizing agent →Cu_2+

Reducing agent → $HCHO$

(4)In fourth reaction oxidized, reduced, oxidizing agent, reducing agent are follows.

Oxidized substance → N_2H_4

Reduced substance → H_2O_2

Oxidizing agent → H_2O_2

Reducing agent → →N_2H_4

(5)In fifth reaction oxidized, reduced oxidizing agent , reducing agent are follows .

(c)Oxidized substance → Pb

Reduced substance → PbO_2

Oxidizing agent → PbO_2

Reducing agent → Pb

Question : 8.14

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

Consider the reaction:

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

$$2S_2O_3^{2-}(aq) + 2Br_2(l) + 5H_2O(l) \rightarrow 2SO_4^{2-}(aq)$$

$$+ 4Br^{\ominus}(aq) + 10H^{\oplus}(aq)$$

Why does the same reducatnt, thiosulphate, react

Answer: Oxidizing agents in the above reaction are $I^2$and Br^2because bromine oxidizes in oxidizing power. Therefore bromine oxidizes Sulphur from +2 to +6 .If we look at the halogens Cl, Br, I, F , then their oxidizing power increases while taking them from the bottom Of these, iodine is less stable than bromine. This is the reason why it oxidizes Safar to +6and Iodine oxidizes it to 2.5

Question 8.15 Justify giving reactions that among halogens,

fluorine is the best oxidant and among compounds, hydro iodic

acid is the best reductant.

Answer: Oxidizing power increases from bottom to top in halogen. I_2am the strongest Oxide from these halogens . After that Cl_2 , Br_2 .We know that fluorine ions are the most stable interest in comparison to other halogens because they can hold (−) charge well.Because these are the most electronegative elements and can be understood by the following reaction.

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

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

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

So F_2 is best oxidizing among .

Hydraulic compound HF, HCl, HBr, HI are reducing agent These are the most reducing agents because only HI, HBr, HCl, HF are ordered for thermal stability. So, HF will be stabilized by hydrogen bonding. If it goes down to group in the hydraulic compound, then the bond length will increase, which will lead to thermal stable increase, due to which it becomes thermally stable. Shown in the reaction below.

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

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

Question 8.16 Why does the following reaction occur?

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

What conclusion about the compound Na_4XeO_6 (of which$XeO^{4-}{}_6$ is a part) can be drawn from the reaction

Answer: F_2 is a very strong oxidizing agent so the reaction F_2toF^-should be very fast because we have strong oxidizing agents but their are $\bar{F}$ to F_2 oxidizing and this is oxidizing by $XeO^{4-}{}_6$ so this tells us that the ones there are very powerful full oxidizing agents. AS this reaction has happened, we can also say that there are Na_4XeO_6 powerful oxidizing agents.

$$\overset{+8}{Xe}O_{6(aq)}^{4-} + 2\overset{-1}{F^-}_{(aq)} + 6H^+_{(aq)} \longrightarrow \overset{+6}{Xe}O_{3(g)} + \overset{0}{F}_{2(g)} + 3H_2O_{(l)}$$

Question:8.17 Consider the reactions.

In the following reactions

(K)

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

(B)

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

(C)

$$C_6H_5CHO(l) + 2\left[Ag(NH_3)_2\right]^+(aq) + 3OH^-(aq) \rightarrow C_6H_6COO^-(aq) + 2Ag(s) + 4NH_3(aq) + 2H_2(l)$$

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

|

From these reactions draw conclusions about the behavior of Tha.

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

Answer: (a) Ag ions are reduced to Ag which is precipitated.

(b)$Cu^2 + +(aq)$ are reduced to Cu which is precipitated.

(C)Ag + (aq) Present in the complex are reduced to Ag Which gets precipitated are shining mirror.

(d)$Cu^2 + (aq)$ ions are not reduced by C^6H^5CHO which is a very weak reducing agent.

So there fore from the above reactions we conclude that Ag + ion is a stronger oxidizing agent then $Cu^2 +$ ion.

Question 8.18 Balance the following redox reactions by ion-electron method:

(a)$MnO^-_4(aq) + \bar{I}(aq) \rightarrow MnO_2(s) + I_2(s)$(in basic medium)

(b)$MnO^-_4(aq) + SO_2(g) \rightarrow Mn^2 + (aq) + HSO^-_4(aq)$(in acidic solutiin)

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

(d)$Cr_2O^{2-}_7 + SO^2(g) \rightarrow Cr^3 + (aq) + SO^{2-}_4(aq)$(in acidic solution)

Answer: in this reaction involved The two half reactions . First one is oxidation half reaction and second one is reduction half reaction .

(a) the reaction is completed in the following steps.

Step (1)Oxidation half reaction –Basically oxidation reaction loses the electron, then this oxidation state is going from -1 to 0.

$$\bar{I}^{1}(aq) \rightarrow I_2(aq)$$

Reduction half reaction – basically in reduction half reaction Mn going from 7 to $+2$ in the reaction then it will gain $3e^{-}$ electrons by reduction

$$M{+}^{7}nO^{-}{}_{4}(aq) \rightarrow M + {}^{4}nO_2(aq)$$

Step (2) In this step we will balance the oxidation reaction and to Balance in R.H.S 2 ē Add Electron.

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

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

Step 3 $-$ In the reduction half reaction of the step, we can see that the Mn is increasing its oxidation state from $+7$ to $+4$. For this, 3 electrons add in the left hand side.

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

After that , to balance the charge, we add 4 OH^{-}ionsions to the right hand side of the reaction as the reaction is taking place in a basic medium.

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

Step 4 – in this step there are 6 O atoms on the right hand side and 4 O atoms on the left hand side. So two water molecules are added to the left hand side.

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

Step 5 – in this step Equalizing the number of electrons by multiplying the oxidation half reaction by 3 and the reduction half reaction by Two

So now ,

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

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

Step 6 – In this step, both half reaction has to be added so that the balance of redox reaction will be as follows

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

(B) Like part a it is a two type reaction, first one is oxidation half reaction and second one is reduction half reaction. Both the reactions are given below.

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

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

Step (b) –In this part the oxidation half reaction is multiplied by 5 and the reduction half reaction is multiplied by 2.After this both add the reaction .By which we get balanced redox reaction. The reaction will get something like this.

$$2MnO_{4(aq)}^{-} + 5SO_{2(g)} + 2H_2O_{(l)} + H^{+}_{(aq)} \longrightarrow 2Mn^{2+}_{(aq)} + 5HSO_{4(aq)}^{-}$$

(C) Like part a it is a two type reaction , first one is oxidation half reaction and second one is a reduction half reaction . Both reaction are given below .

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

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

Step b – in this step Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction . The reaction is given below .

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

(D)Similarly part a it is a two type reaction . The oxidation half reaction are given below.

$$SO_{2(g)} + 2H_2O_{(l)} \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2e^-$$

And reduction half reaction are given below.

$$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e^- \longrightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$$

Step (b) in this step Multiplying the oxidation half reaction by 3 and then adding it to the reduction half reaction, we have the net balanced redox reaction. The reaction are given below.

$$Cr_2O_7^{2-}{}_{(aq)} + 3SO_{2(g)} + 2H^+_{(aq)} \longrightarrow 2Cr^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)} + H_2O_{(l)}$$

Question 8.19 Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidizing agent and the reducing agent.

(a) $P_{4(s)} + OH^-_{(aq)} \longrightarrow PH_{3(g)} + HPO_2^-{}_{(aq)}$

(b) $N_2H_{4(l)} + ClO_3^-{}_{(aq)} \longrightarrow NO_{(g)} + Cl^-_{(g)}$

(c) $Cl_2O_{7(g)} + H_2O_{2(aq)} \longrightarrow ClO_2^-{}_{(aq)} + O_{2(g)} + H^+_{(aq)}$

Answer: (i)We know that the oxidation number of P decrease from 0 in P_4 to -3 in PH_3 and increase from 0 in P_4 to $+2$ in HPO_2.That's why in this reaction P_4 is oxidizing agent .

Ion method equation :The reaction is given below

$$P_{4(s)} \longrightarrow HPO_{2(aq)}^{-}$$

In the above reaction P Atom is balanced. After that reaction is given below

In above reaction oxidation reaction is balanced by adding eight electrons. After that reaction is given

$$\overset{0}{P}_{4(s)} \longrightarrow 4H\overset{2+}{P}O_{2(aq)}^{-}$$

below

$$P_{4(s)} \longrightarrow 4HPO_{2(aq)}^{-} + 8e^{-}$$

In above reaction charged is balance by adding 120 H^{-} After that the reaction is given below

Now the hydrogen and oxygen atoms are balanced by adding $4H_2O$ After that the reaction is given

$$P_{4(s)} + 12OH_{(aq)}^{-} \longrightarrow 4HPO_{2(aq)}^{-} + 8e^{-}$$

below

$$P_{4(s)} + 12OH_{(aq)}^{-} \longrightarrow 4HPO_{2(aq)}^{-} + 4H_2O_{(l)} + 8e^{-} \text{.......(i)}$$

Now the reduction half reaction is given below

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

In the above reaction the P atom is balanced after that reaction is given below

$$\overset{0}{P}_{4(s)} \longrightarrow 4\overset{-3}{P}H_{3(g)}$$

Now the oxidation number is balanced by adding $12e^{-}$ After the reaction is

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

Than the charge is balanced by adding 12 OH^{-} After that the reaction is

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

Now oxygen and hydrogen atoms are balanced by adding $12H_2O$ As the reaction is

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

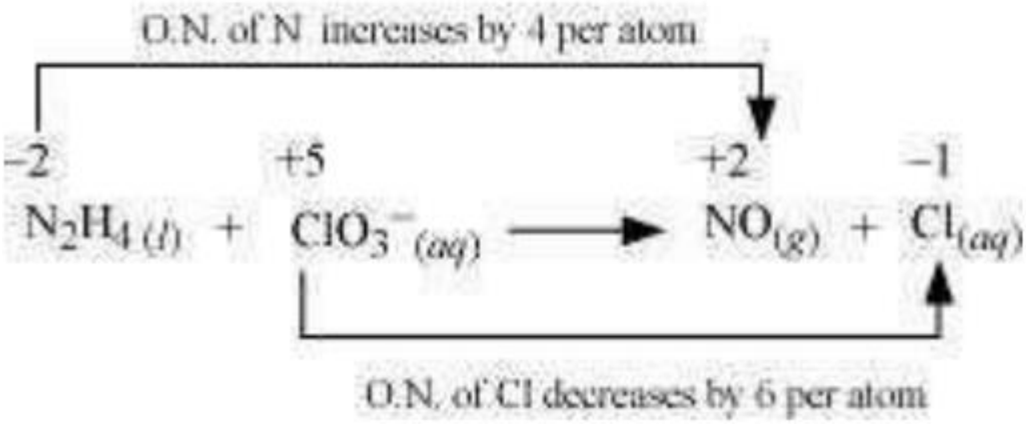
After that in equation first multiplying with 3 and equation second multiplying with 2 and adding both equation. Now we have the balanced chemical equation can be obtained as given

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

(ii)We know that the oxidation number of nitrogen increase from 2 in NH to +2 in No and the oxidation number of Cl is decrease from +5 in to −1 in

Cl^- .That's why in this reaction NH in this reducing agent and is the oxidizing agent .

The reaction is given below



Ion electron method :

In the above reaction N Atom is balanced. After that reaction is given below

In the above reaction N Atom is balanced. After that reaction is given below

$$\overset{-2}{N_2H_{4(l)}} \longrightarrow 2NO_{(g)} \rightarrow \overset{+2}{N}O_{(g)}$$

In above reaction oxidation reaction is balanced by adding eight electrons. After that reaction is given below

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

In above reaction charged is balance by adding $8OH^-$ After that the reaction is given below

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

Now the oxygen atoms are balanced by adding $6H_2O$ After that the reaction is given below

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

Now the reduction half reaction is given below

$$\overset{+5}{Cl}O^-_{3(aq)} \longrightarrow \overset{-1}{Cl}^-_{(aq)}$$

Now the oxidation number is balanced by adding $6e^-$ After the reaction is

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

Than the charge is balanced by adding 6 OH^- After that the reaction is

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

Now oxygen atom are balanced by adding $3H_2O$ As the reaction is

$$ClO_{3(aq)}^{-} + 3H_2O_{(l)} + 6e^{-} \longrightarrow Cl_{(aq)}^{-} + 6OH_{(aq)}^{-}$$

After that in equation first multiplying with 3 and equation second multiplying with 4 and adding both equation. Now we have the balanced chemical equation can be obtained as given

$$3N_2H_{4(l)} + 4ClO_{3(aq)}^{-} \longrightarrow 6NO_{(g)} + 4Cl_{(aq)}^{-} + 6H_2O_{(l)}$$

Oxidizing method:

In above method we get these statements

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

Total increases in oxidation number of $H^2O^2 = 1 \times 6 = 6$

Now multiplying N_2H_4 with 3 and ClO^{3-} with 4 to balance the increase and decrease in the oxidation number so we get this equation .Equation is given below

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

after that the N and Cl atoms are balanced as given below

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

and the O atoms are balanced by adding $6H_2O$

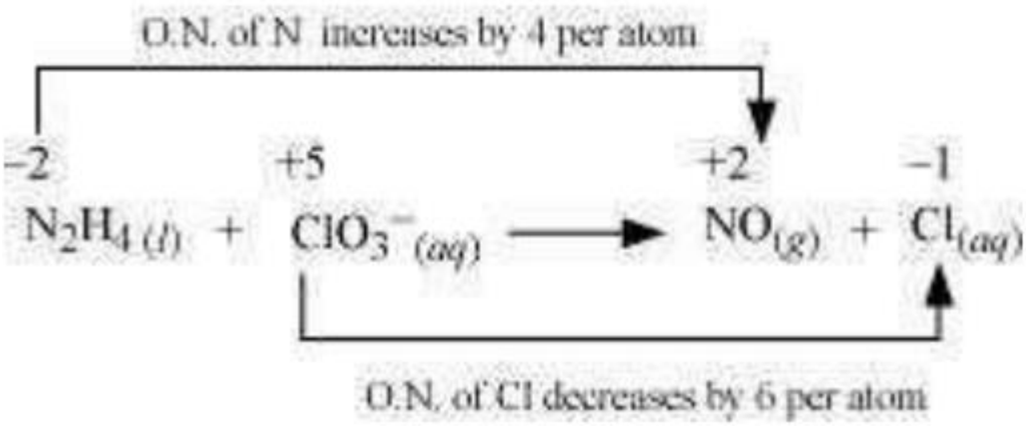
the reaction is given below

$$3N_2H_{4(l)} + 4ClO_{3(aq)}^{-} \longrightarrow 6NO_{(g)} + 4Cl_{(aq)}^{-} + 6H_2O_{(l)}$$

Finally we get the required balanced equation

(iii)We know that the oxidation number of Cl decrease from +7 in Cl_2O_7 to +3 in ClO^{3-} and the oxidation number of Ois increase from -1 in $H_2O_2$to 0 in O_2.That's why in this reaction Cl_2O_7 in this reducing agent and is the oxidizing agentH_2O_2 .

The reaction is given below



Ion electron method :

In the above reaction oxidation half reaction is given below

$$H_2O^{1-}(aq) \rightarrow O_2{}^{0}(g)$$

In above reaction oxidation reaction is balanced by adding two electrons. After that reaction is given below

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

In above reaction charged is balance by adding $2OH^-$ After that the reaction is given below

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

Now the oxygen atoms are balanced by adding $2H_2O$ After that the reaction is given below

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

Now the reduction half reaction is given below

$$\overset{+7}{Cl_2}O_{7(g)} \longrightarrow \overset{+3}{Cl}O^-_{2(aq)}$$

Now the oxidation number is balanced by adding $8e^-$ After the reaction is

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

Than the charge is balanced by adding 6 OH^- After that the reaction is

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

Now oxygen atom are balanced by adding $3H_2O$ As the reaction is

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

After that in equation first multiplying with 4 and adding both equation. Now we have the balanced chemical equation can be obtained as given

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

Oxidizing method:

In above method we get these statements

Total decrease in oxidation number of $Cl_2O_7 = 4 \times 2 = 8$

Total increases in oxidation number of $H_2O_2 = 2 \times 1 = 2$

Now multiplying H_2O_2 and O_2 with 4 to balance the increase and decrease in the oxidation number so we get this equation .

Equation is given below

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

after that the Cl atoms are balanced as given below

and the O atoms are balanced by adding $3H_2O$

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \longrightarrow 2ClO_{2(aq)}^{-} + 4O_{2(g)}$$

the reaction is given below.

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} \longrightarrow 2ClO_{2(aq)}^{-} + 4O_{2(g)} + 3H_2O_{(l)}$$

the hydrogen atoms are balance by adding $2OH^-$

and $2H_2O$.

$$Cl_2O_{7(g)} + 4H_2O_{2(aq)} + 2OH_{(aq)}^{-} \longrightarrow 2ClO_{2(aq)}^{-} + 4O_{2(g)} + 5H_2O_{(l)}$$

Finally we get the required balanced equation

Question 8.20 What sorts of information's can you draw from the following reaction ?

$$(CN)_2(g) + 2O\bar{H}(aq) \rightarrow C\bar{N}(aq) + CN\bar{O}(aq) + H_2O(l)$$

Answer: Oxidation number of Carbon in $(CN)_2$

Oxidation number of N $= -3$

Oxidation number of C $= X$

Now,

$$(X - 3)2 = 0$$

$$2X - 6 = 0$$

$$X = +3$$

Oxidation number of carbon in $C\bar{N}$

Oxidation number of N $= -3$

Lets assume oxidation number of carbon $= X$

Now,

$$X - 3 = -1$$

$$X = +2$$

Oxidation number of carbon in $CN\bar{O}$

Oxidation number of N $= -3$

Oxidation number of O $= -2$

Lets assume oxidation number of carbon $= \mathrm{X}$

Now,

$$\mathrm{X} - 3 - 2 = -1$$

$$\mathrm{X} = +4$$

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

Now oxidation number of carbon in various species

So The oxidation state of carbon is increased(oxidized) and decreased(reduced) as well in the product

$$\left(\overset{+3}{C}N\right)_{2(g)} + 2OH^-_{(aq)} \longrightarrow \overset{+2}{C}N^-_{(aq)} + \overset{+4}{C}NO^-_{(aq)} + H_2O_{(l)}$$

side. So it is a redox reaction and more specifically we can say it disproportion redox reaction.

Question 8.21 The MN^3+ ion is unstable in solution and undergoes disproportionation to give $\mathrm{Mn}^2 + \mathrm{MnO}_2$ and $\mathrm{H}+$ ion. Write a balanced ionic equation for the reaction.

Answer: We can represent the reaction as

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

For this the oxidation half reaction will be something like this

$$\overset{+3}{M}n^{3+}_{(aq)} \longrightarrow \overset{+4}{M}nO_{2(s)}$$

Now we have to add one electron to balance the oxidation reaction so we will get the reaction something like this

To balance the charge, the game 4H + ions will have to be added. So the reaction obtained will be

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

something like this .

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

In the above reaction I have to add $2H_2O$ molecules to balance the O atoms and the H + then the reaction will be something like this .

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

The following are the corresponding reduction half reactions-

One electron has to be added to balance the oxidation number, due to which we will get the following

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

reaction.

To get the balance reaction, we have to add equation (i) and equation (ii) after which we will get the

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

following reaction .

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

Question 8.22 Consider the elements:

F , Cs , I Ne

(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.

Answer: (a)we know that F is the most electronegative element , it means

F Is only element which exhibits only negative oxidation state .

(b)Cs is the Is the most electropositive element , it means Cs is the only element which exhibits only positive oxidation state.

(c)I exhibits both positive and negative oxidation state . I exhibits −1 ,0 , +1 , +3 , +5 , +7 (due to presence of vacant d orbitals) positive oxidation state exhibits of iodine .

(d) Ne Is inert gas , so it is nether positive nor negative oxidation state

Question 8.23 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 equation is

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

The above equation by ion electron method –

(A)First of all write the oxidation and reduction half reaction by observing the change in oxidation number and writing these separately .

Oxidation half reaction are given below –

(B)Balancing the oxidation half reaction –

$$\overset{+4}{S}O_2(aq) \rightarrow \overset{+6}{S}O_4^{2-}(aq)$$

$$\overset{0}{Cl_2}(aq) \rightarrow \overset{-1}{Cl^-}(aq)$$

(i) Add 2 ē Toward right hand side to balance the change on S

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

(ii)Now balance charge by adding 4 H^+Towards left hand side . Reaction is given Below –

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

(iii)Now balance oxygen atoms by adding 2 H^2O Molecules towards left hand side.

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

(C)Now , balance the reduction half reaction –

(i)In reduction half reaction balance $C\bar{l}$ Atoms by multiplying $C\bar{l}$ By 2

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

(ii)Than add electrons towards left hand side to balance the charge . the reaction is given below-

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

(Balanced reduction half reaction)

(D)In last steps adding balanced oxidation half reaction and reduction half reaction . after that the reaction is given below-

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

Hence , this is the balanced redox equation .

Question 8.24 Refer to the periodic table given in your book and no

answer the following questions

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

(b)Select three metals that can show

Answer : we know that about disproportionation reactions, in disproportionation one of the reacting substances always contains an element

And this element can exist in at least three oxidation states. We can understand this by the points given below .

(i)In disproportionation reaction P , Cl , S Can exist three and more oxidation states

(ii)Also in disproportionation reaction Mn , Cu , Ga Are also exists in three and more oxidation states.

Question 8.25 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 with10.00 g of ammonia and 20.00 g of oxygen?

Answer : we know that ,

Mass of ammonia $= 17$ g

Mass of oxygen molecule $= 32$ g

Mass of nitric oxide $= 30$ g

Mass of water $= 18$ g

The reaction unvalued in the manufacturing process is

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

So now,

$$4 \times 17 = 68 \text{ g} \quad , \qquad 5 \times 37 = 160 \text{ g} \, , \qquad 4 \times 30 = 120 \text{ g}$$

Now,

From the available data – 68g Of NH^3 Will react with $O^2 = 160$ g

10 g of NH^3Will react with $O^2 =$

$$10 \text{ g of } NH_3 \text{ will react with } O_2 = \frac{(160 \text{ g})}{(68 \text{ g})} \times (10 \text{ g}) = 23.6 \text{ g}$$

But oxygen which is actually available (20.09) Is less than the amount which is needed . There fore oxygen is the limiting reactant

Now,160 g of O_2 Will from NO = 120 g

$$\therefore 20 \text{ g of } O_2 \text{ will form NO} = \frac{(120 \text{ g})}{(106 \text{ g})} \times (20 \text{ g})$$
$$= 15 \text{ g}$$

Question 8.26 Using the standard electrode potentials given in the Table predict if the reaction between the following is feasible:

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

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

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

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

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

Answer : We have the possible reaction between $Fe^3 +$ and I^-

Now,

$$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$$ is given by,

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

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

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

The values ofE^0in the above reaction are positions. Thus the reactions between Fe^3 +and I^- are feasible.

(b)Similarly we have the possible reaction between $Ag +$ and Cu

Now,

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

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

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

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

If value of E^0 Is positive than above reaction is positive so , the reaction between Ag + and Cu Are feasible.

(c)Similarly we have the possible reaction between $Fe^3 + (aq)$ and $Cu(s)$

Now ,

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

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

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

So E^0 is Positive than above reaction is positive . finally the reaction between $Fe^3 + (aq)$and $Cu(s)$ are feasible.

(d) Similarly here , we have the possible reaction between $Ag(s)$ and $Fe^3 + (aq)$

Now , the reaction is given below

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

is given by,

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

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

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

So E^0 is negative that’s why above reaction is negative , hence the reaction between $Ag(s)$ and $Fe^3 + (aq)$ are not feasible .

(e)Now we have the possible reaction between $Br_2(aq)$ and $Fe^2+(aq)$ .the reaction is given below

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

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

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

So E^0 is positive that's why above reaction is positive ,hence the reaction between $Br_2(aq)$and $Fe^2+(aq)$ are feasible.

Question 8.27 Predict the products of electrolysis in each of the following:

(i) An aqueous solution of $AgNO_3$with silver electrodes

(ii)An aqueous solution $AgNO_3$ with platinum electrodes

(iii)A dilute solution of H_2SO_4 with platinum electrodes

(iv) An aqueous solution of $CuCl_2$with platinum electrodes.

Answer : In aqueous solution of $AgNO^3$with silver electrodes , we can see
At cathode -Silver ions have lower discharge potential than hydrogen ions . So ,silver ions will be deposited in preference to hydrogen ions and
At anode –Silver anode will dissolve to form silver ions in the solution.
The reactions are given below.

$$Ag \rightarrow Ag^+ + \bar{e}$$

(ii) in aqueous solution of $AgNO_3$with platinum electrodes , we can see
At cathode – Silver ions have lower discharge potential than hydrogen ions. So , silver ions will be deposited in preference to hydrogen ions.
At anode –Hydroxide ions having lower discharge potential will be discharged in preference to nitrate ions. Hydroxide ions will decompose to give oxygen. The reaction for this is given below

$$4O\bar{H}(aq) \rightarrow 2H_2O + O_2(g) + 4e^-$$

(iii) A dilute solution of H_2SO_4 with platinum electrodes. We can se
At cathode - $2H^+ + 2\bar{e} \rightarrow H_2(g)$
At anode – we know that the Hydroxide ions having lower discharge potential will be discharged in preference to sulphate ions. Hydroxide ions will decompose to give oxygen. The reaction is given below

$$4O\bar{H}(aq) \rightarrow 2H^2O(l) + O^{2(g)} + 4e$$

(iv) In aqueous solution of $CuCl_2$ with platinum electrodes we can see
At cathode - Cupric ions will be reduced in preference to protons .Reaction is given below

$$Cu_2 + 2\bar{e} \rightarrow Cu$$

at anode - Chloride ions will be oxidized in preference to hydroxide ions

$$2C\bar{l} \rightarrow Cl^2 + 2e^-$$

Question 8.28 Arrange the following metals in the order in which they displace

each other from the solution of their salts.

Al , Cu , Fe , Mg , Zn

Answer: we know that the order in which metals displace each other from the solution of their salts can be given with the help of their standard electrode potential. Since magnesium has the least standard electrode potential so it is the most strong reducing agent. So the required order we get is given below

Mg › Al ›Zn › Fe › Cu

Question8.29 Given the standard electrode potentials $K+/K = -2.93V, Ag+/Ag = 0.80V$, $Hg^2+/Hg = 0.79\ V$, $Mg^2+/Mg = -2.37\ V$, $Cr^3+/Cr = -0.74$

Arrange these metals in their increasing order of reducing power.

Answer : we already know that lower the reduction potential and the higher is the reducing power. The given standard electrode potentials increase in the order of

$$K+/K < Mg^2+/Mg < Cr^3+/Cr < Hg^2+/Hg < Ag+/Ag$$

Hence, In this way it is proved that the reducing power of the given metals increases $Ag < Hg < Cr < Mg < K$ in the following order are given below .

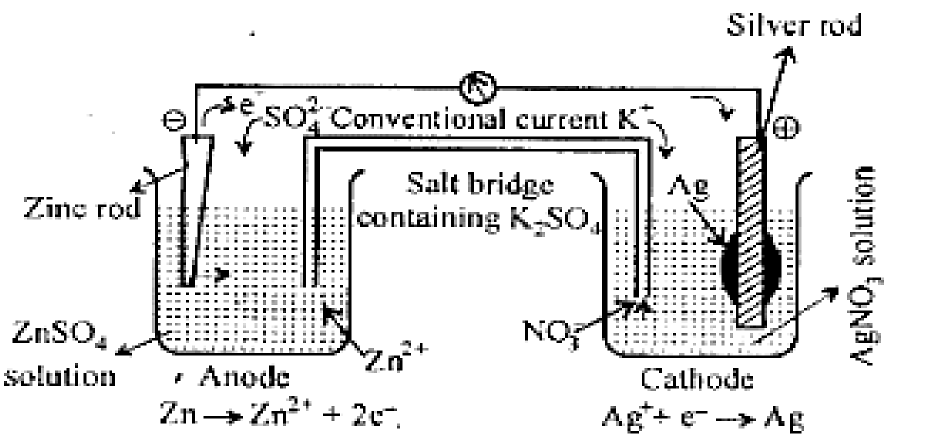
Question 8.30 Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq)) \rightarrow Zn^2 + (aq) + 2Ag(s)$ takes place, further show:

(i)which of the electrode is negatively charged,

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

(iii) individual reaction at each electrode

Answer :



The cell will be represented as:

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

We know that the galvanic cell in which the given reaction takes place is depicted as given below .

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

(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,

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

The reaction taking place at the cathode is given by,

$$\mathrm{Ag^{+}(aq)} + \bar{e} \rightarrow \mathrm{Ag(s)}$$