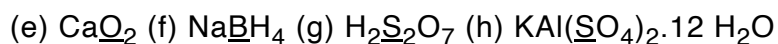
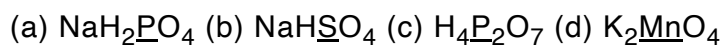


## Chapter 1 – Redox Reactions

Page No 272:

Question 8.1:

Assign oxidation numbers to the underlined elements in each of the following species:



Answer:



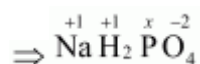
Let the oxidation number of P be x.

We know that,

Oxidation number of Na = +1

Oxidation number of H = +1

Oxidation number of O = -2



Then, we have

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

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

$$\Rightarrow x = +5$$

Hence, the oxidation number of P is +5.



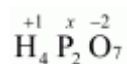
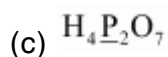
Then, we have

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

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

$$\Rightarrow x = +6$$

Hence, the oxidation number of S is +6.



Then, we have

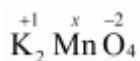
$$4(+1) + 2(x) + 7(-2) = 0$$

$$\Rightarrow 4 + 2x - 14 = 0$$

$$\Rightarrow 2x = +10$$

$$\Rightarrow x = +5$$

Hence, the oxidation number of P is + 5.



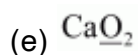
Then, we have

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

$$\Rightarrow 2 + x - 8 = 0$$

$$\Rightarrow x = +6$$

Hence, the oxidation number of Mn is + 6.



Then, we have

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

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

$$\Rightarrow x = -1$$

Hence, the oxidation number of O is – 1.



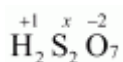
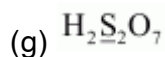
Then, we have

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

$$\Rightarrow 1 + x - 4 = 0$$

$$\Rightarrow x = +3$$

Hence, the oxidation number of B is + 3.



Then, we have

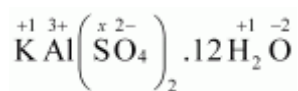
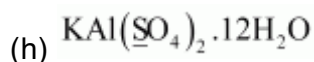
$$2(+1) + 2(x) + 7(-2) = 0$$

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

$$\Rightarrow 2x = 12$$

$$\Rightarrow x = +6$$

Hence, the oxidation number of S is + 6.



Then, we have

$$\begin{aligned} 1(+1) + 1(+3) + 2(x) + 8(-2) + 24(+1) + 12(-2) &= 0 \\ \Rightarrow 1 + 3 + 2x - 16 + 24 - 24 &= 0 \\ \Rightarrow 2x &= 12 \\ \Rightarrow x &= +6 \end{aligned}$$

Or,

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

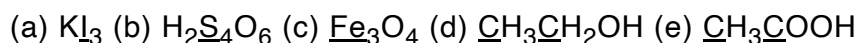
$$\begin{aligned} 1(+1) + 1(+3) + 2(x) + 8(-2) &= 0 \\ \Rightarrow 1 + 3 + 2x - 16 &= 0 \\ \Rightarrow 2x &= 12 \\ \Rightarrow x &= +6 \end{aligned}$$

Hence, the oxidation number of S is + 6.

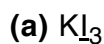
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Question 8.2:

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



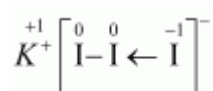
Answer:



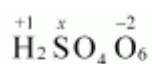
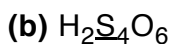
In  $\text{KI}_3$ , the oxidation number (O.N.) of K is +1. Hence, the average oxidation number of I

is  $-\frac{1}{3}$ . However, O.N. cannot be fractional. Therefore, we will have to consider the structure of  $\text{KI}_3$  to find the oxidation states.

In a  $\text{KI}_3$  molecule, an atom of iodine forms a coordinate covalent bond with an iodine molecule.



Hence, in a  $\text{KI}_3$  molecule, the O.N. of the two I atoms forming the  $\text{I}_2$  molecule is 0, whereas the O.N. of the I atom forming the coordinate bond is +1.



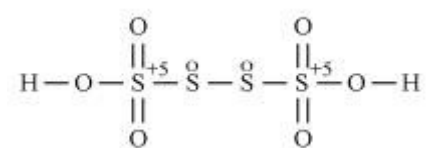
$$\text{Now, } 2(+1) + 4(x) + 6(-2) = 0$$

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

$$\Rightarrow 4x = 10$$

$$\Rightarrow x = +2\frac{1}{2}$$

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

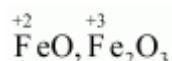


The O.N. of two of the four S atoms is +5 and the O.N. of the other two S atoms is 0.



On taking the O.N. of O as  $\hat{=}$  -2, the O.N. of Fe is found to be  $+2\frac{2}{3}$ . However, O.N. cannot be fractional.

Here, one of the three Fe atoms exhibits the O.N. of +2 and the other two Fe atoms exhibit the O.N. of +3.



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

$$\text{or, } 2x + 4 = 0$$

$$\text{or, } x = -2$$

Hence, the O.N. of C is  $\hat{=}$  -2.

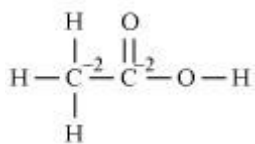


$$2(x) + 4(+1) + 2(-2) = 0$$

$$\text{or, } 2x = 0$$

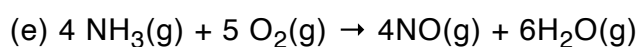
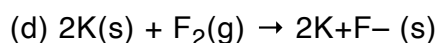
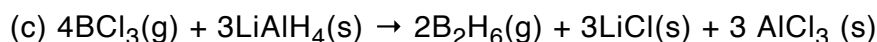
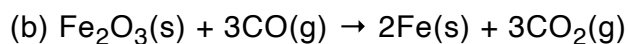
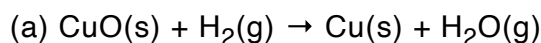
$$\text{or, } x = 0$$

However, 0 is average O.N. of C. The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number. Thus, C exhibits the oxidation states of +2 and  $\hat{=}$  -2 in  $\text{CH}_3\text{COOH}$ .

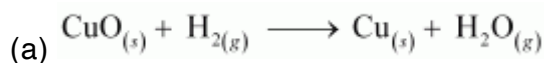


Question 8.3:

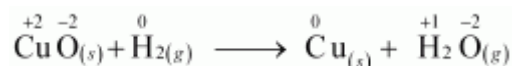
Justify that the following reactions are redox reactions:



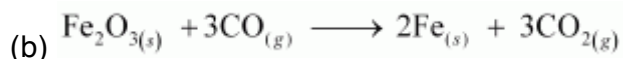
Answer:



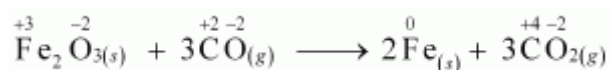
Let us write the oxidation number of each element involved in the given reaction as:



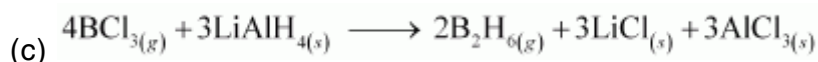
Here, the oxidation number of Cu decreases from +2 in CuO to 0 in Cu i.e., CuO is reduced to Cu. Also, the oxidation number of H increases from 0 in H<sub>2</sub> to +1 in H<sub>2</sub>O i.e., H<sub>2</sub> is oxidized to H<sub>2</sub>O. Hence, this reaction is a redox reaction.



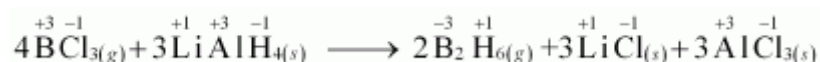
Let us write the oxidation number of each element in the given reaction as:



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

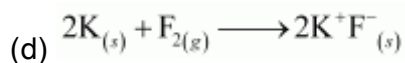


The oxidation number of each element in the given reaction can be represented as:

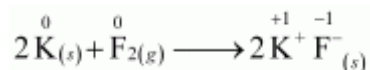


In this reaction, the oxidation number of B decreases from +3 in BCl<sub>3</sub> to -3 in B<sub>2</sub>H<sub>6</sub>. i.e., BCl<sub>3</sub> is reduced to B<sub>2</sub>H<sub>6</sub>. Also, the oxidation number of H increases from -1 in LiAlH<sub>4</sub> to +1

in  $B_2H_6$  i.e.,  $LiAlH_4$  is oxidized to  $B_2H_6$ . Hence, the given reaction is a redox reaction.

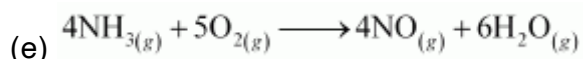


The oxidation number of each element in the given reaction can be represented as:

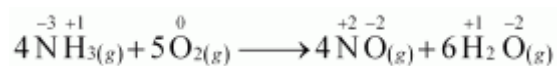


In this reaction, the oxidation number of K increases from 0 in K to +1 in KF i.e., K is oxidized to KF. On the other hand, the oxidation number of F decreases from 0 in  $F_2$  to -1 in KF i.e.,  $F_2$  is reduced to KF.

Hence, the above reaction is a redox reaction.



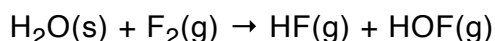
The oxidation number of each element in the given reaction can be represented as:



Here, the oxidation number of N increases from -3 in  $NH_3$  to +2 in NO. On the other hand, the oxidation number of  $O_2$  decreases from 0 in  $O_2$  to -2 in NO and  $H_2O$  i.e.,  $O_2$  is reduced. Hence, the given reaction is a redox reaction.

Question 8.4:

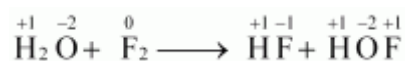
Fluorine reacts with ice and results in the change:



Justify that this reaction is a redox reaction.

Answer:

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



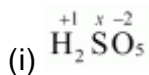
Here, we have observed that the oxidation number of F increases from 0 in  $F_2$  to +1 in HOF. Also, the oxidation number decreases from 0 in  $F_2$  to -1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

Question 8.5:

Calculate the oxidation number of sulphur, chromium and nitrogen in  $H_2SO_5$ ,

$Cr_2O_7^{2-}$  and  $NO_3^-$ . Suggest structure of these compounds. Count for the fallacy.

Answer:



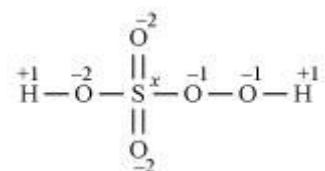
$$2(+1) + 1(x) + 5(-2) = 0$$

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

$$\Rightarrow x = +8$$

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  $\text{H}_2\text{SO}_5$  is shown as follows:

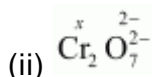


$$\text{Now, } 2(+1) + 1(x) + 3(-2) + 2(-1) = 0$$

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

$$\Rightarrow x = +6$$

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



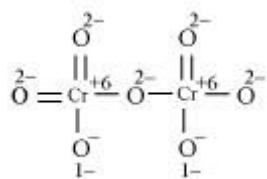
$$2(x) + 7(-2) = -2$$

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

$$\Rightarrow x = +6$$

Here, there is no fallacy about the O.N. of Cr in  $\text{Cr}_2\text{O}_7^{2-}$ .

The structure of  $\text{Cr}_2\text{O}_7^{2-}$  is shown as follows:



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

(iii)



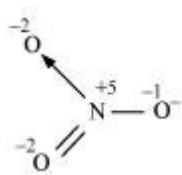
$$1(x) + 3(-2) = -1$$

$$\Rightarrow x - 6 = -1$$

$$\Rightarrow x = +5$$

Here, there is no fallacy about the O.N. of N in  $\text{NO}_3^-$ .

The structure of  $\text{NO}_3^-$  is shown as follows:



The N atom exhibits the O.N. of +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(III) oxide  |

Answer:

(a) Mercury (II) chloride:



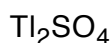
(b) Nickel (II) sulphate:



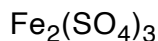
(c) Tin (IV) oxide:



(d) Thallium (I) sulphate:



(e) Iron (III) sulphate:



(f) Chromium (III) oxide:



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

---



Substance	O.N. of carbon
$\text{CH}_2\text{Cl}_2$	0
$\text{ClC}\equiv\text{CCl}$	+1
$\text{HC}\equiv\text{CH}$	-1
$\text{CHCl}_3, \text{CO}$	+2
$\text{CH}_3\text{Cl}$	-2
$\text{Cl}_3\text{C} - \text{CCl}_3$	+3
$\text{H}_3\text{C} - \text{CH}_3$	-3
$\text{CCl}_4, \text{CO}_2$	+4
$\text{CH}_4$	-4

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

Substance	O.N. of nitrogen
$\text{N}_2$	0
$\text{N}_2\text{O}$	+1
$\text{N}_2\text{H}_2$	-1
$\text{NO}$	+2
$\text{N}_2\text{H}_4$	-2
$\text{N}_2\text{O}_3$	+3
$\text{NH}_3$	-3
$\text{NO}_2$	+4



+5

Question 8.8:

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

Answer:

In sulphur dioxide ( $\text{SO}_2$ ), the oxidation number (O.N.) of S is +4 and the range of the O.N. that S can have is from +6 to -2.

Therefore,  $\text{SO}_2$  can act as an oxidising as well as a reducing agent.

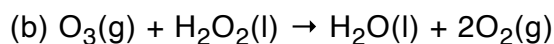
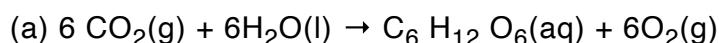
In hydrogen peroxide ( $\text{H}_2\text{O}_2$ ), the O.N. of O is -1 and the range of the O.N. that O can have is from 0 to -2. O can sometimes also attain the oxidation numbers +1 and +2. Hence,  $\text{H}_2\text{O}_2$  can act as an oxidising as well as a reducing agent.

In ozone ( $\text{O}_3$ ), the O.N. of O is zero and the range of the O.N. that O can have is from 0 to -2. Therefore, the O.N. of O can only decrease in this case. Hence,  $\text{O}_3$  acts only as an oxidant.

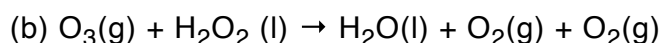
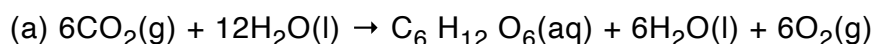
In nitric acid ( $\text{HNO}_3$ ), the O.N. of N is +5 and the range of the O.N. that N can have is from +5 to -3. Therefore, the O.N. of N can only decrease in this case. Hence,  $\text{HNO}_3$  acts only as an oxidant.

Question 8.9:

Consider the reactions:



Why it is more appropriate to write these reactions as:



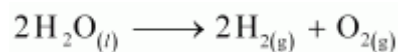
Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

Answer:

(a) The process of photosynthesis involves two steps.

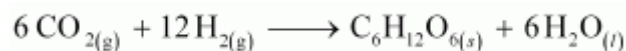
Step 1:

H<sub>2</sub>O decomposes to give H<sub>2</sub> and O<sub>2</sub>.

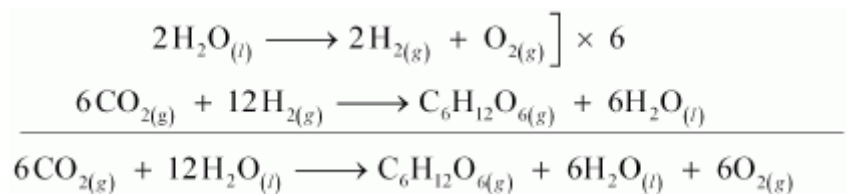


Step 2:

The H<sub>2</sub> produced in step 1 reduces CO<sub>2</sub>, thereby producing glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>) and H<sub>2</sub>O.



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

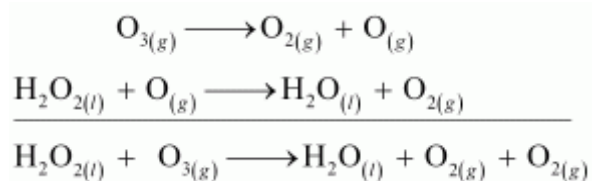


It is more appropriate to write the reaction as given above because water molecules are also produced in the process of photosynthesis.

The path of this reaction can be investigated by using radioactive H<sub>2</sub>O<sup>18</sup> in place of H<sub>2</sub>O.

(b) O<sub>2</sub> is produced from each of the two reactants O<sub>3</sub> and H<sub>2</sub>O<sub>2</sub>. For this reason, O<sub>2</sub> is written twice.

The given reaction involves two steps. First, O<sub>3</sub> decomposes to form O<sub>2</sub> and O. In the second step, H<sub>2</sub>O<sub>2</sub> reacts with the O produced in the first step, thereby producing H<sub>2</sub>O and O<sub>2</sub>.



The path of this reaction can be investigated by using <sup>18</sup>H<sub>2</sub>O<sub>2</sub> or <sup>18</sup>O<sub>3</sub>.

Question 8.10:

The compound AgF<sub>2</sub> is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

Answer:

The oxidation state of Ag in AgF<sub>2</sub> is +2. But, +2 is an unstable oxidation state of Ag.

Therefore, whenever AgF<sub>2</sub> is formed, silver readily accepts an electron to form Ag<sup>+</sup>. This

helps to bring the oxidation state of Ag down from +2 to a more stable state of +1. As a result,  $\text{AgF}_2$  acts as a very strong oxidizing agent.

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Question 8.11:

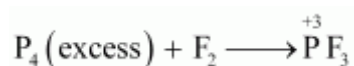
Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

Answer:

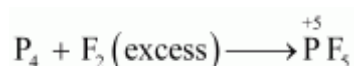
Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. This can be illustrated as follows:

(i)  $\text{P}_4$  and  $\text{F}_2$  are reducing and oxidising agents respectively.

If an excess of  $\text{P}_4$  is treated with  $\text{F}_2$ , then  $\text{PF}_3$  will be produced, wherein the oxidation number (O.N.) of P is +3.

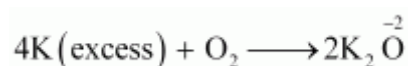


However, if  $\text{P}_4$  is treated with an excess of  $\text{F}_2$ , then  $\text{PF}_5$  will be produced, wherein the O.N. of P is +5.

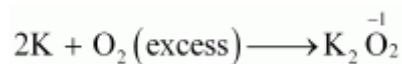


(ii) K acts as a reducing agent, whereas  $\text{O}_2$  is an oxidising agent.

If an excess of K reacts with  $\text{O}_2$ , then  $\text{K}_2\text{O}$  will be formed, wherein the O.N. of O is -2.

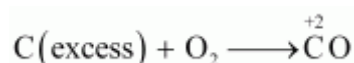


However, if K reacts with an excess of  $\text{O}_2$ , then  $\text{K}_2\text{O}_2$  will be formed, wherein the O.N. of O is -1.



(iii) C is a reducing agent, while  $\text{O}_2$  acts as an oxidising agent.

If an excess of C is burnt in the presence of insufficient amount of  $\text{O}_2$ , then CO will be produced, wherein the O.N. of C is +2.



On the other hand, if C is burnt in an excess of  $\text{O}_2$ , then  $\text{CO}_2$  will be produced, wherein the O.N. of C is +4.



Question 8.12:

How do you count for the following observations?

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

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

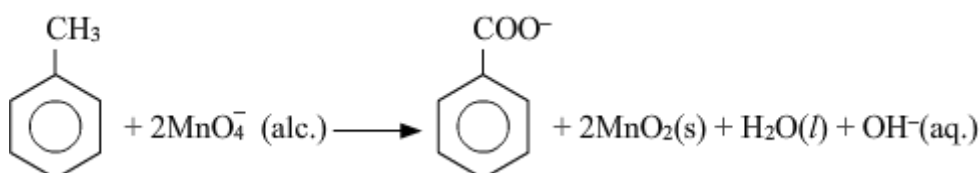
Answer:

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

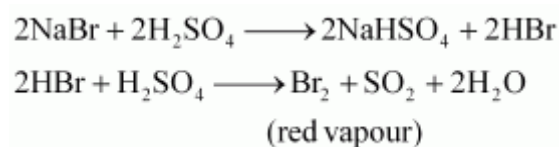
(i) In a neutral medium,  $\text{OH}^-$  ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii)  $\text{KMnO}_4$  and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol,  $\text{KMnO}_4$  and toluene can react at a faster rate.

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



(b) When conc.  $\text{H}_2\text{SO}_4$  is added to an inorganic mixture containing bromide, initially HBr is produced. HBr, being a strong reducing agent reduces  $\text{H}_2\text{SO}_4$  to  $\text{SO}_2$  with the evolution of red vapour of bromine.

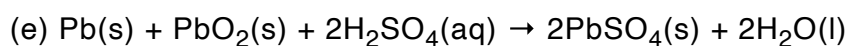
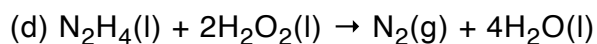
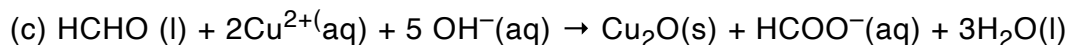
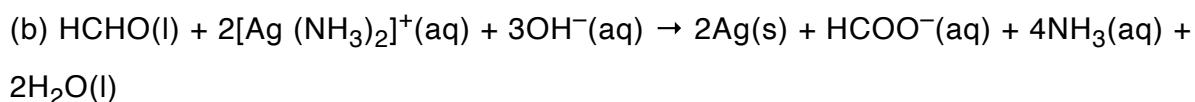


But, when conc.  $\text{H}_2\text{SO}_4$  is added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, being a weak reducing agent, cannot reduce  $\text{H}_2\text{SO}_4$  to  $\text{SO}_2$ .



Question 8.13:

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



Answer:

(a) Oxidised substance  $\rightarrow \text{C}_6\text{H}_6\text{O}_2$

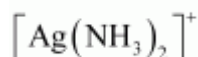
Reduced substance  $\rightarrow \text{AgBr}$

Oxidising agent  $\rightarrow \text{AgBr}$

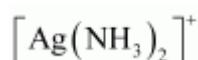
Reducing agent  $\rightarrow \text{C}_6\text{H}_6\text{O}_2$

(b) Oxidised substance  $\rightarrow \text{HCHO}$

Reduced substance  $\rightarrow$



Oxidising agent  $\rightarrow$



Reducing agent  $\rightarrow \text{HCHO}$

(c) Oxidised substance  $\rightarrow \text{HCHO}$

Reduced substance  $\rightarrow \text{Cu}^{2+}$

Oxidising agent  $\rightarrow \text{Cu}^{2+}$

Reducing agent  $\rightarrow \text{HCHO}$

(d) Oxidised substance  $\rightarrow \text{N}_2\text{H}_4$

Reduced substance  $\rightarrow \text{H}_2\text{O}_2$

Oxidising agent  $\rightarrow \text{H}_2\text{O}_2$

Reducing agent  $\rightarrow \text{N}_2\text{H}_4$

(e) Oxidised substance  $\rightarrow \text{Pb}$

Reduced substance  $\rightarrow \text{PbO}_2$

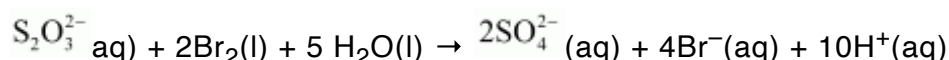
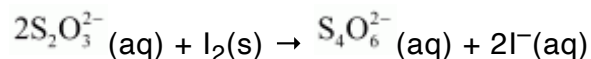
Oxidising agent  $\rightarrow \text{PbO}_2$

Reducing agent  $\rightarrow \text{Pb}$

---

Question 8.14:

Consider the reactions:



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

Answer:

The average oxidation number (O.N.) of S in  $\text{S}_2\text{O}_3^{2-}$  is +2. Being a stronger oxidising agent than  $\text{I}_2$ ,  $\text{Br}_2$  oxidises  $\text{S}_2\text{O}_3^{2-}$  to  $\text{SO}_4^{2-}$ , in which the O.N. of S is +6. However,  $\text{I}_2$  is a weak oxidising agent. Therefore, it oxidises  $\text{S}_2\text{O}_3^{2-}$  to

$\text{S}_4\text{O}_6^{2-}$ , in which the average O.N. of S is only +2.5. As a result,  $\text{S}_2\text{O}_3^{2-}$  reacts differently with iodine and bromine.

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Question 8.15:

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

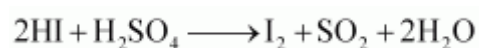
Answer:

$\text{F}_2$  can oxidize  $\text{Cl}^-$  to  $\text{Cl}_2$ ,  $\text{Br}^-$  to  $\text{Br}_2$ , and  $\text{I}^-$  to  $\text{I}_2$  as:



On the other hand,  $\text{Cl}_2$ ,  $\text{Br}_2$ , and  $\text{I}_2$  cannot oxidize  $\text{F}^-$  to  $\text{F}_2$ . The oxidizing power of halogens increases in the order of  $\text{I}_2 < \text{Br}_2 < \text{Cl}_2 < \text{F}_2$ . Hence, fluorine is the best oxidant among halogens.

$\text{HI}$  and  $\text{HBr}$  can reduce  $\text{H}_2\text{SO}_4$  to  $\text{SO}_2$ , but  $\text{HCl}$  and  $\text{HF}$  cannot. Therefore,  $\text{HI}$  and  $\text{HBr}$  are stronger reductants than  $\text{HCl}$  and  $\text{HF}$ .



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



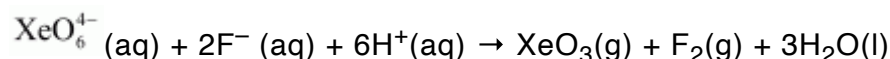
Hence, hydroiodic acid is the best reductant among hydrohalic compounds.

Thus, the reducing power of hydrohalic acids increases in the order of  $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$ .

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Question 8.16:

Why does the following reaction occur?

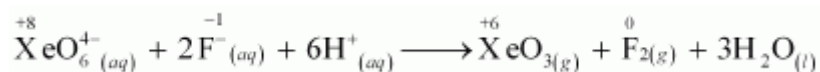


What conclusion about the compound  $\text{Na}_4\text{XeO}_6$  (of which  $\text{XeO}_6^{4-}$  is a part) can be drawn from the reaction.

Answer:

The given reaction occurs because  $\text{XeO}_6^{4-}$  oxidises

$\text{F}^-$  and  $\text{F}^-$  reduces  $\text{XeO}_6^{4-}$ .



In this reaction, the oxidation number (O.N.) of Xe decreases from +8 in

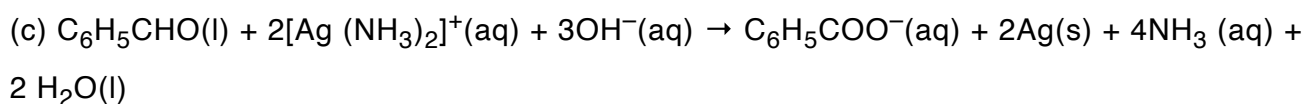
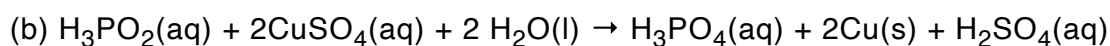
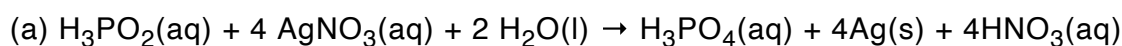
$\text{XeO}_6^{4-}$  to +6 in  $\text{XeO}_3$  and the O.N. of F increases from -1 in  $\text{F}^-$  to 0 in  $\text{F}_2$ .

Hence, we can conclude that  $\text{Na}_4\text{XeO}_6$  is a stronger oxidising agent than  $\text{F}^-$ .

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Question 8.17:

Consider the reactions:



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

Answer:



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

In reaction (c),  $\text{Ag}^+$  oxidises  $\text{C}_6\text{H}_5\text{CHO}$  to  $\text{C}_6\text{H}_5\text{COO}^-$ , but in reaction (d),  $\text{Cu}^{2+}$  cannot oxidise  $\text{C}_6\text{H}_5\text{CHO}$ .

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

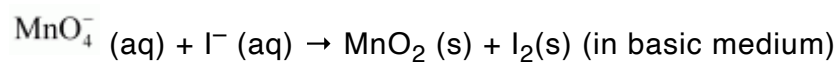
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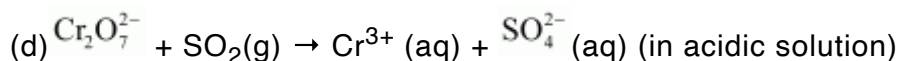
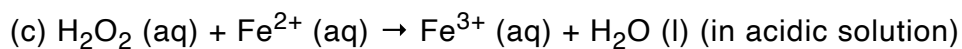
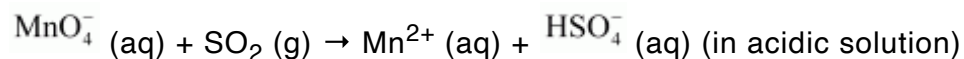
Question 8.18:

Balance the following redox reactions by ion-electron method:

(a)

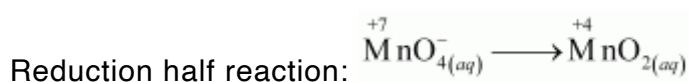
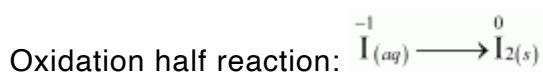


(b)



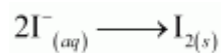
Answer:

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

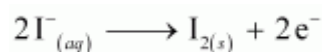


Step 2:

Balancing I in the oxidation half reaction, we have:

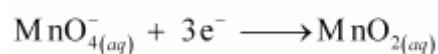


Now, to balance the charge, we add  $2\text{e}^-$  to the RHS of the reaction.

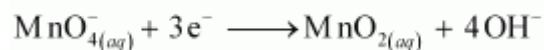


Step 3:

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

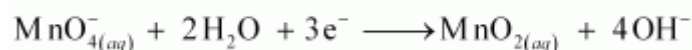


Now, to balance the charge, we add 4 OH<sup>-</sup> ions to the RHS of the reaction as the reaction is taking place in a basic medium.



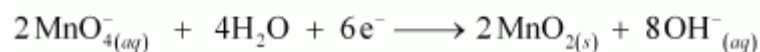
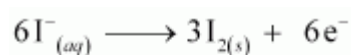
Step 4:

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



Step 5:

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



Step 6:

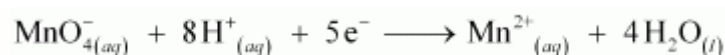
Adding the two half reactions, we have the net balanced redox reaction as:



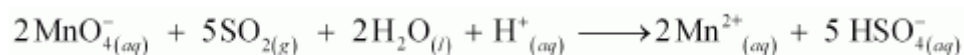
(b) Following the steps as in part (a), we have the oxidation half reaction as:



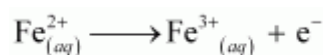
And the reduction half reaction as:



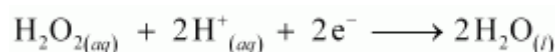
Multiplying the oxidation half reaction by 5 and the reduction half reaction by 2, and then by adding them, we have the net balanced redox reaction as:



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



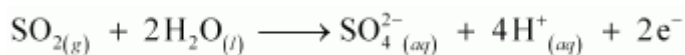
And the reduction half reaction as:



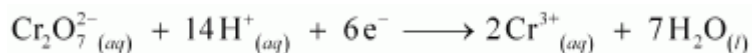
Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:



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



And the reduction half reaction as:



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



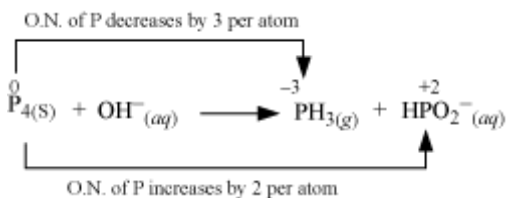
Question 8.19:

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



Answer:

(a)



The O.N. (oxidation number) of P decreases from 0 in  $\text{P}_4$  to  $-3$  in  $\text{PH}_3$  and increases from 0 in  $\text{P}_4$  to  $+2$  in  $\text{HPO}_2^-$ . Hence,  $\text{P}_4$  acts both as an oxidising agent and a reducing agent in this reaction.

**Ion-electron method:**

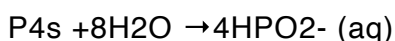
The oxidation half equation is:



The P atom is balanced as:



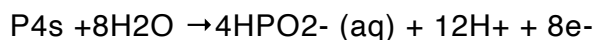
The O atom is balanced by adding 8  $\text{H}_2\text{O}$  molecules:



The H atom is balanced by adding 12  $\text{H}^+$  ions:

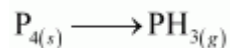


The charge is balanced by adding  $\text{e}^-$  as:

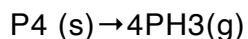


...(i)

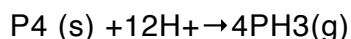
The reduction half equation is:



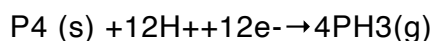
The P atom is balanced as:



The H is balanced by adding 12 H<sup>+</sup> as:



The charge is balanced by adding 12e<sup>-</sup> as:

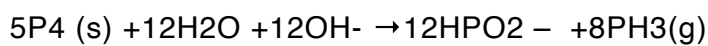


...(ii)

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



As, the medium is basic, add 12OH<sup>-</sup> both sides as:



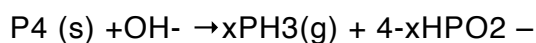
This is the required balanced equation.

### **Oxidation number method:**

Let, total no of P reduced = x

∴

Total no of P oxidised = 4 - x



... (i) Total decrease in oxidation number of P = x × 3 = 3x

Total increase in oxidation number of P = (4 - x) × 2 = 8 - 2x

∴

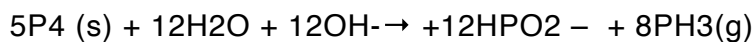
$$3x = 8 - 2x \quad x = 8/5 \text{ From (i),}$$



Since, reaction occurs in basic medium, the charge is balanced by adding 7OH<sup>-</sup> on LHS as:

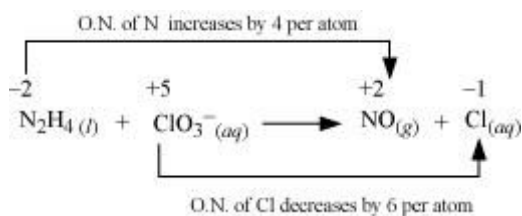


The O atoms are balanced by adding 12H<sub>2</sub>O as:



This is the required balanced equation.

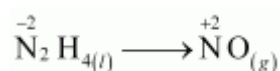
(b)



The oxidation number of N increases from  $-2$  in  $\text{N}_2\text{H}_4$  to  $+2$  in  $\text{NO}$  and the oxidation number of Cl decreases from  $+5$  in  $\text{ClO}_3^-$  to  $-1$  in  $\text{Cl}^-$ . Hence, in this reaction,  $\text{N}_2\text{H}_4$  is the reducing agent and  $\text{ClO}_3^-$  is the oxidizing agent.

**Ion-electron method:**

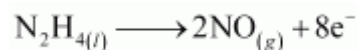
The oxidation half equation is:



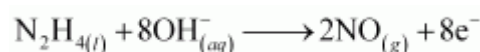
The N atoms are balanced as:



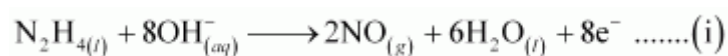
The oxidation number is balanced by adding 8 electrons as:



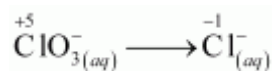
The charge is balanced by adding 8  $\text{OH}^-$  ions as:



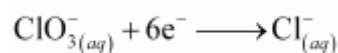
The O atoms are balanced by adding 6  $\text{H}_2\text{O}$  as:



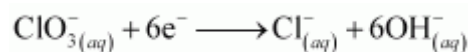
The reduction half equation is:



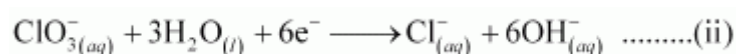
The oxidation number is balanced by adding 6 electrons as:



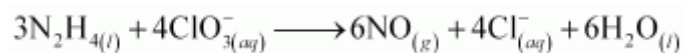
The charge is balanced by adding 6  $\text{OH}^-$  ions as:



The O atoms are balanced by adding 3  $\text{H}_2\text{O}$  as:



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

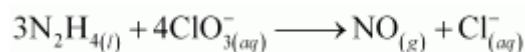


### Oxidation number method:

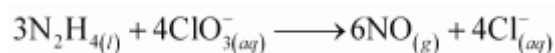
Total decrease in oxidation number of N =  $2 \times 4 = 8$

Total increase in oxidation number of Cl =  $1 \times 6 = 6$

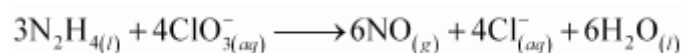
On multiplying  $\text{N}_2\text{H}_4$  with 3 and  $\text{ClO}_3^-$  with 4 to balance the increase and decrease in O.N., we get:



The N and Cl atoms are balanced as:

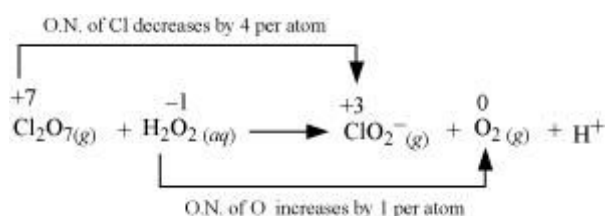


The O atoms are balanced by adding  $6\text{H}_2\text{O}$  as:



This is the required balanced equation.

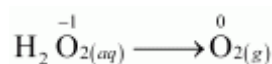
### (c)



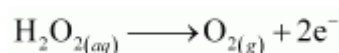
The oxidation number of Cl decreases from + 7 in  $\text{Cl}_2\text{O}_7$  to + 3 in  $\text{ClO}_2^-$  and the oxidation number of O increases from - 1 in  $\text{H}_2\text{O}_2$  to zero in  $\text{O}_2$ . Hence, in this reaction,  $\text{Cl}_2\text{O}_7$  is the oxidizing agent and  $\text{H}_2\text{O}_2$  is the reducing agent.

### Ion–electron method:

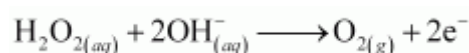
The oxidation half equation is:



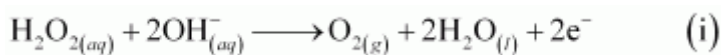
The oxidation number is balanced by adding 2 electrons as:



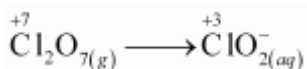
The charge is balanced by adding  $2\text{OH}^-$  ions as:



The oxygen atoms are balanced by adding  $2\text{H}_2\text{O}$  as:



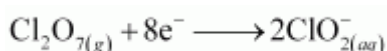
The reduction half equation is:



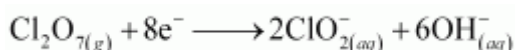
The Cl atoms are balanced as:



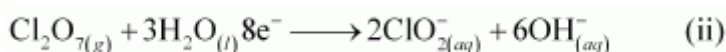
The oxidation number is balanced by adding 8 electrons as:



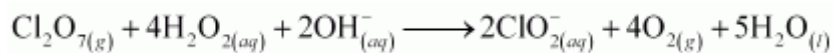
The charge is balanced by adding 6OH<sup>-</sup> as:



The oxygen atoms are balanced by adding 3H<sub>2</sub>O as:



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

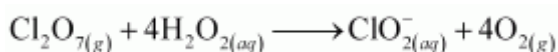


#### Oxidation number method:

Total decrease in oxidation number of Cl<sub>2</sub>O<sub>7</sub> = 4 × 2 = 8

Total increase in oxidation number of H<sub>2</sub>O<sub>2</sub> = 2 × 1 = 2

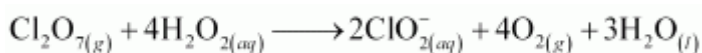
By multiplying H<sub>2</sub>O<sub>2</sub> and O<sub>2</sub> with 4 to balance the increase and decrease in the oxidation number, we get:



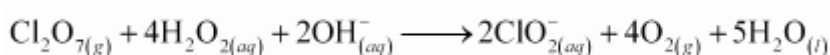
The Cl atoms are balanced as:



The O atoms are balanced by adding 3H<sub>2</sub>O as:



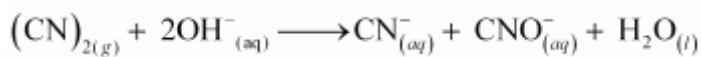
The H atoms are balanced by adding 2OH<sup>-</sup> and 2H<sub>2</sub>O as:



This is the required balanced equation.

Question 8.20:

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



Answer:

The oxidation numbers of carbon in  $(\text{CN})_2$ ,  $\text{CN}^-$  and  $\text{CNO}^-$  are +3, +2 and +4 respectively. These are obtained as shown below:

Let the oxidation number of C be  $x$ .



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

$$\therefore x = 3$$



$$x - 3 = -1$$

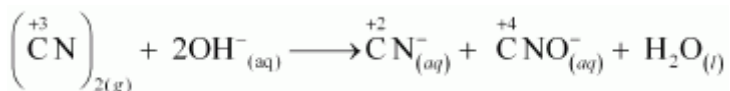
$$\therefore x = 2$$



$$x - 3 - 2 = -1$$

$$\therefore x = 4$$

The oxidation number of carbon in the various species is:



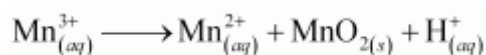
It can be easily observed that the same compound is being reduced and oxidised simultaneously in the given equation. Reactions in which the same compound is reduced and oxidised is known as disproportionation reactions. Thus, it can be said that the alkaline decomposition of cyanogen is an example of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

Question 8.21:

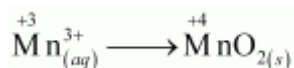
The  $\text{Mn}^{3+}$  ion is unstable in solution and undergoes disproportionation to give  $\text{Mn}^{2+}$ ,  $\text{MnO}_2$ , and  $\text{H}^+$  ion. Write a balanced ionic equation for the reaction.

Answer:

The given reaction can be represented as:

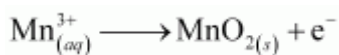


The oxidation half equation is:

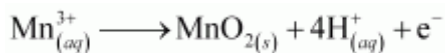


The oxidation number is balanced by adding one electron as:





The charge is balanced by adding  $4\text{H}^{+}$  ions as:



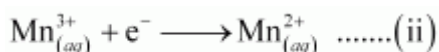
The O atoms and  $\text{H}^{+}$  ions are balanced by adding  $2\text{H}_2\text{O}$  molecules as:



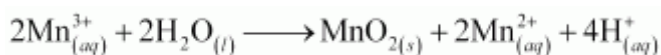
The reduction half equation is:



The oxidation number is balanced by adding one electron as:



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



Question 8.22:

Consider the elements:

Cs, Ne, I and F

- Identify the element that exhibits only negative oxidation state.
- Identify the element that exhibits only positive oxidation state.
- Identify the element that exhibits both positive and negative oxidation states.
- Identify the element which exhibits neither the negative nor does the positive oxidation state.

Answer:

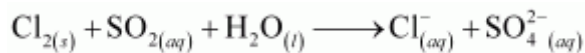
- F exhibits only negative oxidation state of  $-1$ .
  - Cs exhibits positive oxidation state of  $+1$ .
  - I exhibits both positive and negative oxidation states. It exhibits oxidation states of  $-1$ ,  $+1$ ,  $+3$ ,  $+5$ , and  $+7$ .
  - The oxidation state of Ne is zero. It exhibits neither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

Question 8.23:

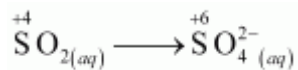
Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Answer:

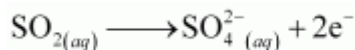
The given redox reaction can be represented as:



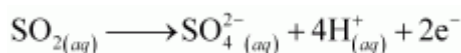
The oxidation half reaction is:



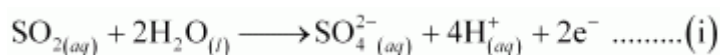
The oxidation number is balanced by adding two electrons as:



The charge is balanced by adding  $4\text{H}^+$  ions as:



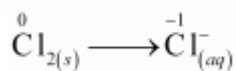
The O atoms and  $\text{H}^+$  ions are balanced by adding  $2\text{H}_2\text{O}$  molecules as:



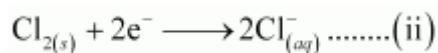
The reduction half reaction is:



The chlorine atoms are balanced as:



The oxidation number is balanced by adding electrons



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



Question 8.24:

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

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

Answer:

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

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

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

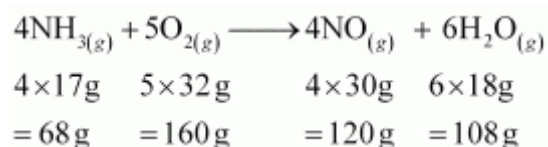
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Question 8.25:

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 with 10.00 g. of ammonia and 20.00 g of oxygen?

Answer:

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



Thus, 68 g of  $\text{NH}_3$  reacts with 160 g of  $\text{O}_2$ .

Therefore, 10g of  $\text{NH}_3$  reacts with  $\frac{160 \times 10}{68}$  g of  $\text{O}_2$ , or 23.53 g of  $\text{O}_2$ .

But the available amount of  $\text{O}_2$  is 20 g.

Therefore,  $\text{O}_2$  is the limiting reagent (we have considered the amount of  $\text{O}_2$  to calculate the weight of nitric oxide obtained in the reaction).

Now, 160 g of  $\text{O}_2$  gives 120g of NO.

Therefore, 20 g of  $\text{O}_2$  gives  $\frac{120 \times 20}{160}$  g of N, or 15 g of NO.

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

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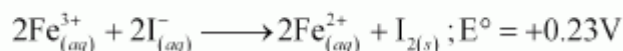
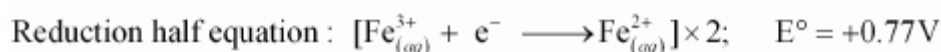
Question 8.26:

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

- (a)  $\text{Fe}^{3+}(\text{aq})$  and  $\text{I}^{-}(\text{aq})$
- (b)  $\text{Ag}^{+}(\text{aq})$  and  $\text{Cu}(\text{s})$
- (c)  $\text{Fe}^{3+}(\text{aq})$  and  $\text{Cu}(\text{s})$
- (d)  $\text{Ag}(\text{s})$  and  $\text{Fe}^{3+}(\text{aq})$
- (e)  $\text{Br}_2(\text{aq})$  and  $\text{Fe}^{2+}(\text{aq})$

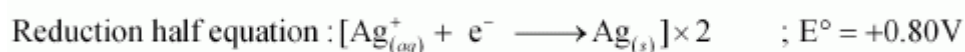
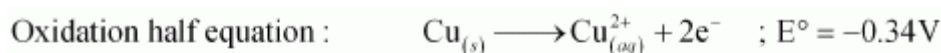
Answer:

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



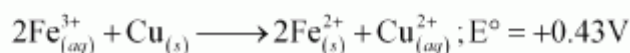
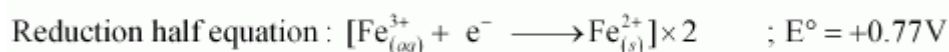
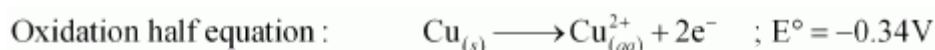
$E^{\circ}$  for the overall reaction is positive. Thus, the reaction between  $\text{Fe}^{3+}_{(aq)}$  and  $\text{I}^{-}_{(aq)}$  is feasible.

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



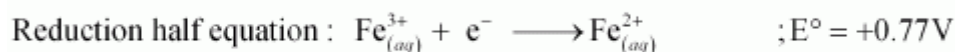
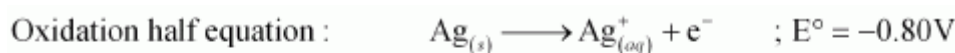
$E^{\circ}$  positive for the overall reaction is positive. Hence, the reaction between  $\text{Ag}^{+}_{(aq)}$  and  $\text{Cu}_{(s)}$  is feasible.

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



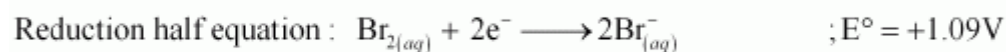
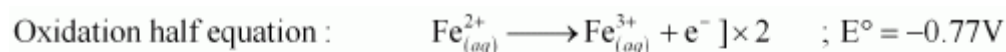
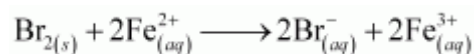
$E^{\circ}$  positive for the overall reaction is positive. Hence, the reaction between  $\text{Fe}^{3+}_{(aq)}$  and  $\text{Cu}_{(s)}$  is feasible.

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



Here,  $E^\circ$  for the overall reaction is negative. Hence, the reaction between  $\text{Ag}_{(s)}$  and  $\text{Fe}_{(aq)}^{3+}$  is not feasible.

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



Here,  $E^\circ$  for the overall reaction is positive. Hence, the reaction between  $\text{Br}_{2(aq)}$  and  $\text{Fe}_{(aq)}^{2+}$  is feasible.

Page No 275:

Question 8.27:

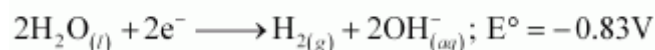
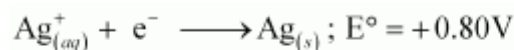
Predict the products of electrolysis in each of the following:

- (i) An aqueous solution of  $\text{AgNO}_3$  with silver electrodes
- (ii) An aqueous solution  $\text{AgNO}_3$  with platinum electrodes
- (iii) A dilute solution of  $\text{H}_2\text{SO}_4$  with platinum electrodes
- (iv) An aqueous solution of  $\text{CuCl}_2$  with platinum electrodes.

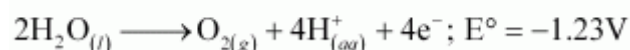
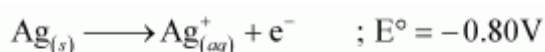
Answer:

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

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



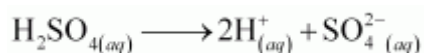
Hence,  $\text{Ag}^+$  ions are reduced at the cathode. Similarly,  $\text{Ag}$  metal or  $\text{H}_2\text{O}$  molecules can be oxidized at the anode. But the oxidation potential of  $\text{Ag}$  is higher than that of  $\text{H}_2\text{O}$  molecules.



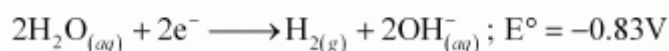
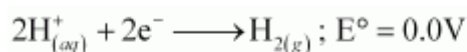
Therefore,  $\text{Ag}$  metal gets oxidized at the anode.

(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.



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.



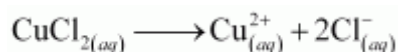
Hence, at the cathode,  $H^+$  ions are reduced to liberate  $H_2$  gas.

On the other hand, at the anode, either of  $SO_4^{2-}$  ions or  $H_2O$  molecules can get oxidized.

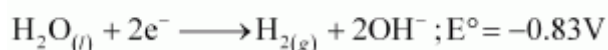
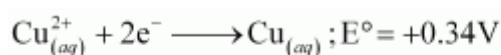
But the oxidation of  $SO_4^{2-}$  involves breaking of more bonds than that of  $H_2O$  molecules.

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:

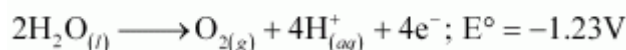
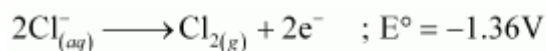


On electrolysis, either of  $Cu^{2+}$  ions or  $H_2O$  molecules can get reduced at the cathode. But the reduction potential of  $Cu^{2+}$  is more than that of  $H_2O$  molecules.



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

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



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.

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

Al, Cu, Fe, Mg and Zn.

Answer:

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

The order of the increasing reducing power of the given metals is  $\text{Cu} < \text{Fe} < \text{Zn} < \text{Al} < \text{Mg}$ .

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:

$\text{Mg} > \text{Al} > \text{Zn} > \text{Fe} > \text{Cu}$

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Question 8.29:

Given the standard electrode potentials,

$\text{K}^+/\text{K} = -2.93\text{V}$ ,  $\text{Ag}^+/\text{Ag} = 0.80\text{V}$ ,

$\text{Hg}^{2+}/\text{Hg} = 0.79\text{V}$

$\text{Mg}^{2+}/\text{Mg} = -2.37\text{V}$ .  $\text{Cr}^{3+}/\text{Cr} = -0.74\text{V}$

Arrange these metals in their increasing order of reducing power.

Answer:

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

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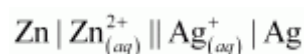
Question 8.30:

Depict the galvanic cell in which the reaction  $\text{Zn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$  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:

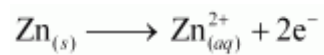
The galvanic cell corresponding to the given redox reaction can be represented as:



**(i)** Zn electrode is negatively charged because at this electrode, Zn oxidizes to  $\text{Zn}^{2+}$  and the leaving electrons accumulate on this electrode.

**(ii)** Ions are the carriers of current in the cell.

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



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

