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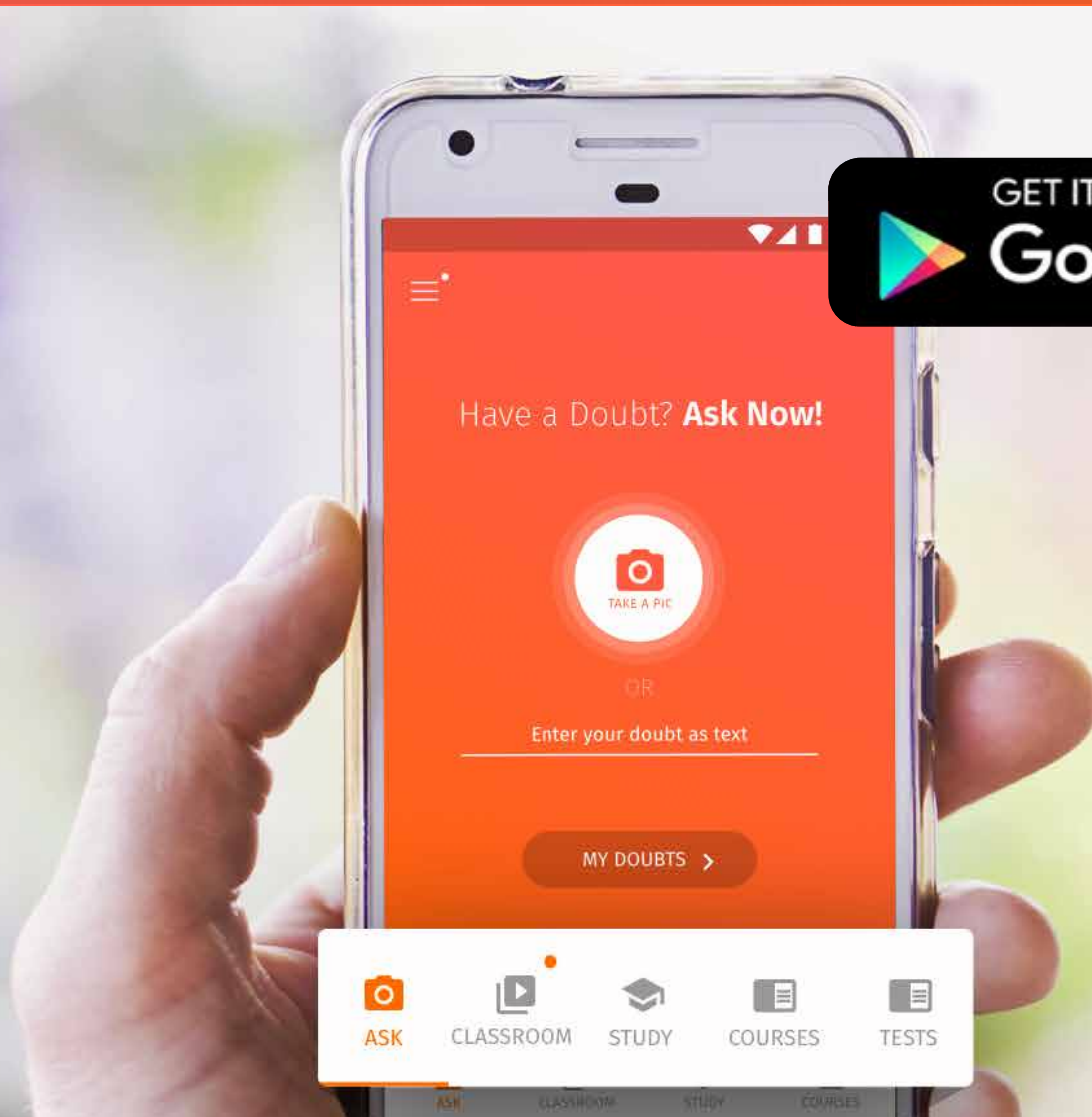
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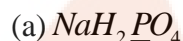
NCERT Exercise

Question 1:

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

- (a) $\text{NaH}_2\text{P}\underline{\text{O}}_4$ (b) $\text{NaH}\underline{\text{S}}\text{O}_4$ (c) $\text{H}_4\underline{\text{P}}_2\text{O}_7$ (d) $\text{K}_2\underline{\text{Mn}}\text{O}_4$
 (e) $\text{Ca}\underline{\text{O}}_2$ (f) $\text{Na}\underline{\text{B}}\text{H}_4$ (g) $\text{H}_2\underline{\text{S}}_2\text{O}_7$ (h) $\text{KAl}(\underline{\text{S}}\text{O}_4)_2 \cdot 12 \text{H}_2\text{O}$

Solution 1:



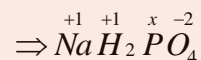
Let the oxidation number of P be x .

We know that,

Oxidation number of Na = +1

Oxidation number of H = +1

Oxidation number of O = -2



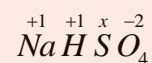
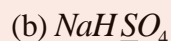
Then, we have

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

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

$$\Rightarrow x = +5$$

Hence, the oxidation number of P is +5.



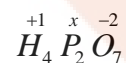
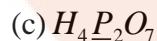
Then, we have

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

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

$$\Rightarrow x = +6$$

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.

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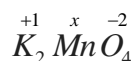
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(d) $K_2\overset{+1}{\text{K}}\overset{x}{\text{Mn}}\overset{-2}{\text{O}}_4$ 

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.

(e) $\text{Ca}\overset{+2}{\text{Ca}}\overset{x}{\text{O}}_2$ 

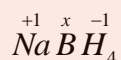
Then, we have

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

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

$$\Rightarrow x = -1$$

Hence, the oxidation number of O is – 1.

(f) $\text{Na}\overset{+1}{\text{Na}}\overset{x}{\text{B}}\overset{-1}{\text{H}}_4$ 

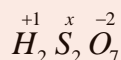
Then, we have

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

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

$$\Rightarrow x = +3$$

Hence, the oxidation number of B is + 3.

(g) $\text{H}_2\overset{+1}{\text{H}}_2\overset{x}{\text{S}}_2\overset{-2}{\text{O}}_7$ 

Then, we have

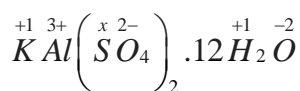
$$2(+1) + 2(x) + 7(-2) = 0$$

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

$$\Rightarrow 2x = 12$$

$$\Rightarrow x = +6$$

Hence, the oxidation number of S is + 6.

(h) $\text{KAl}(\overset{x}{\text{S}}\overset{2-}{\text{O}}_4)_2 \cdot 12\text{H}_2\overset{+1}{\text{H}}_2\overset{-2}{\text{O}}$ 

Then, we have

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

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

$$\Rightarrow 2x = 12$$

$$\Rightarrow x = +6$$

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

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

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

$$\Rightarrow 2x = 12$$

$$\Rightarrow x = +6$$

Hence, the oxidation number of S is + 6.

Question 2:

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

- (a) $K\underline{I}_3$ (b) $H_2\underline{S}_4O_6$ (c) \underline{Fe}_3O_4 (d) $\underline{C}H_3\underline{C}H_2OH$ (e) $\underline{C}H_3\underline{C}OOH$

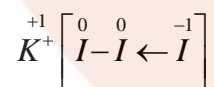
Solution 2:

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

In 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

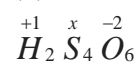
KI_3 to find the oxidation states.

In a KI_3 molecule, an atom of iodine forms a coordinate covalent bond with an iodine molecule.



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

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



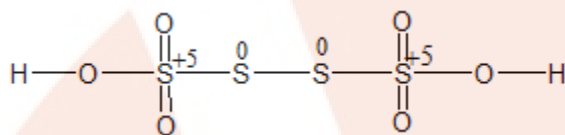
$$\text{Now, } 2(+1) + 4(x) + 6(-2) = 0$$

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

$$\Rightarrow 4x = 10$$

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

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

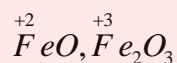


The O.N. of two of the four S atoms is +5 and the O.N. of the other two S atoms is 0.

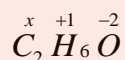
(c) Fe_3O_4

On taking the O.N. of O as -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.



(d) $\text{CH}_3\text{CH}_2\text{OH}$



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

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

$$x = -2$$

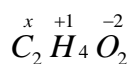
The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number.

Thus, C exhibits the oxidation states of -3 and -1



Ethanol

(e) CH_3COOH



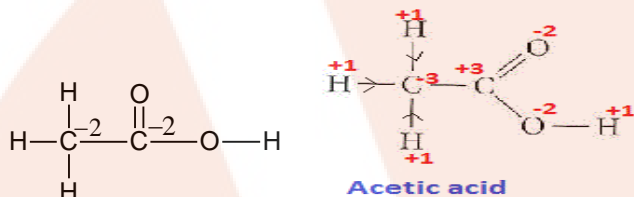
$$2(x) + 4(+1) + 2(-2) = 0$$

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

$$\Rightarrow x = 0$$

However, 0 is average O.N. of C. The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number.

Thus, C exhibits the oxidation states of + 3 and – 3 in CH_3COOH .

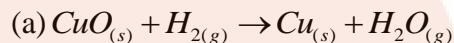


Question 3:

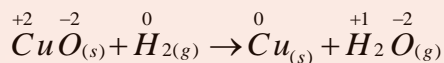
Justify that the following reactions are redox reactions:

- $\text{CuO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)}$
- $\text{Fe}_2\text{O}_3\text{(s)} + 3\text{CO(g)} \rightarrow 2\text{Fe(s)} + 3\text{CO}_2\text{(g)}$
- $4\text{BCl}_3\text{(g)} + 3\text{LiAlH}_4\text{(s)} \rightarrow 2\text{B}_2\text{H}_6\text{(g)} + 3\text{LiCl(s)} + 3\text{AlCl}_3\text{(s)}$
- $2\text{K(s)} + \text{F}_2\text{(g)} \rightarrow 2\text{K}^+\text{F}^-\text{(s)}$
- $4\text{NH}_3\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 4\text{NO(g)} + 6\text{H}_2\text{O(g)}$

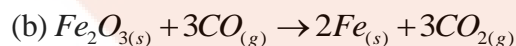
Solution 3:



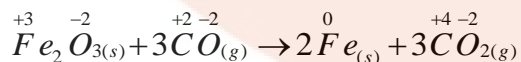
Let us write the oxidation number of each element involved in the given reaction as:



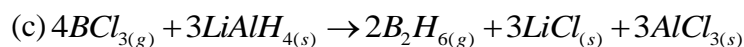
Here, the oxidation number of Cu decreases from +2 in CuO to 0 in Cu i.e., CuO is reduced to Cu. Also, the oxidation number of H increases from 0 in H_2 to +1 in H_2O i.e., H_2 is oxidized to H_2O . Hence, this reaction is a redox reaction.



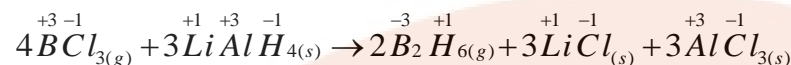
Let us write the oxidation number of each element in the given reaction as:



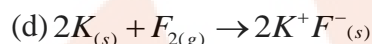
Here, the oxidation number of Fe decreases from +3 in Fe_2O_3 to 0 in Fe i.e., Fe_2O_3 is reduced to Fe. On the other hand, the oxidation number of C increases from +2 in CO to +4 in CO_2 i.e., CO is oxidized to CO_2 . Hence, the given reaction is a redox reaction.



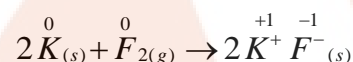
The oxidation number of each element in the given reaction can be represented as:



In this reaction, the oxidation number of B decreases from +3 in BCl_3 to -3 in B_2H_6 . i.e., BCl_3 is reduced to B_2H_6 . Also, the oxidation number of H increases from -1 in $LiAlH_4$ to +1 in B_2H_6 i.e., $LiAlH_4$ is oxidized to B_2H_6 . Hence, the given reaction is a redox reaction.

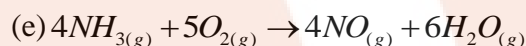


The oxidation number of each element in the given reaction can be represented as:

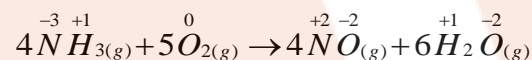


In this reaction, the oxidation number of K increases from 0 in K to +1 in KF i.e., K is oxidized to KF. On the other hand, the oxidation number of F decreases from 0 in F_2 to -1 in KF i.e., F_2 is reduced to KF.

Hence, the above reaction is a redox reaction.



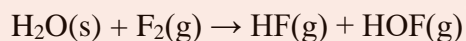
The oxidation number of each element in the given reaction can be represented as:



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 4:

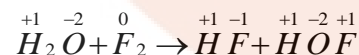
Fluorine reacts with ice and results in the change:



Justify that this reaction is a redox reaction.

Solution 4:

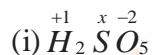
Let us write the oxidation number of each atom involved in the given reaction above its symbol as:



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 5:

Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_5 , $\text{Cr}_2\text{O}_7^{2-}$ and NO_3^- . Suggest structure of these compounds. Count for the fallacy.

Solution 5:

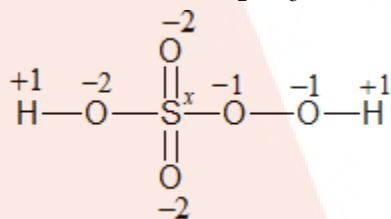
$$2(+1) + 1(x) + 5(-2) = 0$$

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

$$\Rightarrow x = +8$$

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

The structure of H_2SO_5 is shown as follows:



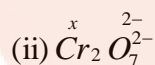
$$2(\text{H}) + 1(\text{S}) + 3(\text{O}) + 2(\text{O in peroxy linkage})$$

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

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

$$\Rightarrow x = +6$$

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



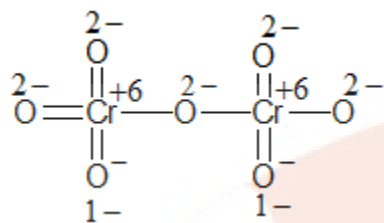
$$2(x) + 7(-2) = -2$$

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

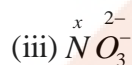
$$\Rightarrow x = +6$$

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.



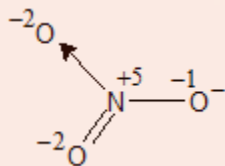
$$1(x) + 3(-2) = -1$$

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

$$\Rightarrow x = +5$$

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

The structure of NO_3^- is shown as follows:



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

Question 6:

Write the formulae for the following compounds:

- Mercury(II) chloride
- Nickel(II) sulphate
- Tin(IV) oxide
- Thallium(I) sulphate
- Iron(III) sulphate
- Chromium(III) oxide

Solution 6:

(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 7:

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

Solution 7:

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

Substance	O.N. of carbon
CH_2Cl_2	0
$\text{ClC}\equiv\text{CCl}$	+1
$\text{HC}\equiv\text{CH}$	-1
CHCl_3, CO	+2
CH_3Cl	-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
CH_4	-4

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

Substance	O.N. of nitrogen
N_2	0
N_2O	+1
N_2H_2	-1
NO	+2
N_2H_4	-2
N_2O_3	+3
NH_3	-3
NO_2	+4
N_2O_5	+5

Question 8:

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

Solution 8:

In sulphur dioxide (SO_2), the oxidation number (O.N.) of S is +4 and the range of the O.N. that S can have is from +6 to -2.

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

In hydrogen peroxide (H_2O_2), the O.N. of O is -1 and the range of the O.N. that O can have is from 0 to -2. O can sometimes also attain the oxidation numbers +1 and +2.

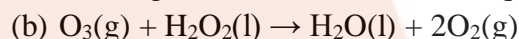
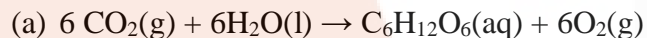
Hence, H_2O_2 can act as an oxidising as well as a reducing agent.

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

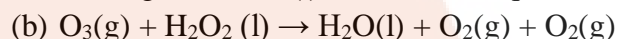
In nitric acid (HNO_3), the O.N. of N is +5 and the range of the O.N. that N can have is from +5 to -3. Therefore, the O.N. of N can only decrease in this case. Hence, HNO_3 acts only as an oxidant.

Question 9:

Consider the reactions:



Why it is more appropriate to write these reactions as:



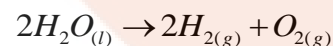
Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

Solution 9:

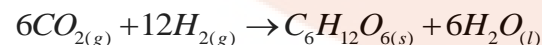
(a) The process of photosynthesis involves two steps.

Step 1:

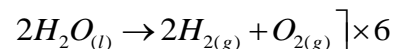
H_2O decomposes to give H_2 and O_2 .

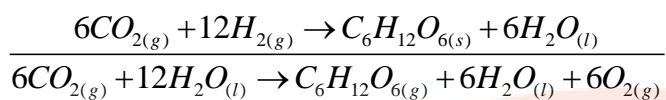
**Step 2:**

The H_2 produced in **step 1** reduces CO_2 , thereby producing glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) and H_2O .



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



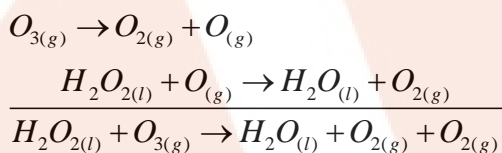


It is more appropriate to write the reaction as given above because water molecules are also produced in the process of photosynthesis.

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

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

The given reaction involves two steps. First, O_3 decomposes to form O_2 and O . In the second step, H_2O_2 reacts with the O produced in the first step, thereby producing H_2O and O_2 .



The path of this reaction can be investigated by using $\text{H}_2\text{O}_2^{18}$ or O_3^{18} .

Question 10:

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

Solution 10:

The oxidation state of Ag in AgF_2 is +2. But, +2 is an unstable oxidation state of Ag . Therefore, whenever AgF_2 is formed, silver readily accepts an electron to form Ag^+ . This helps to bring the oxidation state of Ag down from +2 to a more stable state of +1. As a result, AgF_2 acts as a very strong oxidizing agent.

Question 11:

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

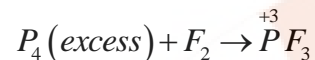
Solution 11:

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. This can be

illustrated as follows:

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

If an excess of P_4 is treated with F_2 , then PF_3 will be produced, wherein the oxidation number (O.N.) of P is +3.

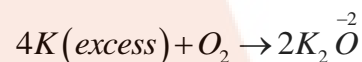


However, if P_4 is treated with an excess of F_2 , then PF_5 will be produced, wherein the O.N. of P is +5.

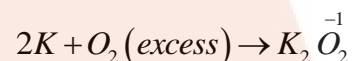


(ii) K acts as a reducing agent, whereas O_2 is an oxidising agent.

If an excess of K reacts with O_2 , then K_2O will be formed, wherein the O.N. of O is -2.

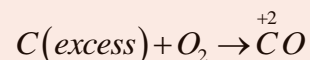


However, if K reacts with an excess of O_2 , then K_2O_2 will be formed, wherein the O.N. of O is -1.

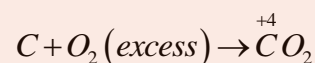


(iii) C is a reducing agent, while O_2 acts as an oxidising agent.

If an excess of C is burnt in the presence of insufficient amount of 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 O_2 , then CO_2 will be produced, wherein the O.N. of C is +4.



Question 12:

How do you count for the following observations?

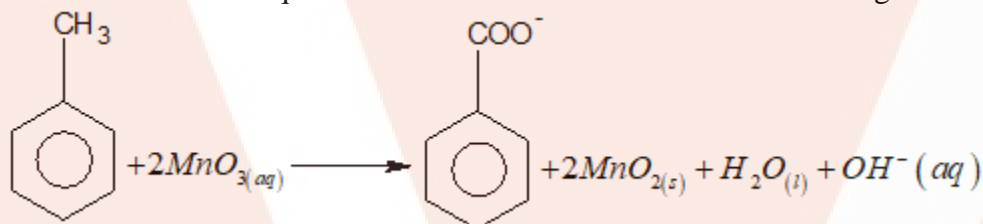
- Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.
- When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl , but if the mixture contains bromide then we get red vapour of bromine. Why?

Solution 12:

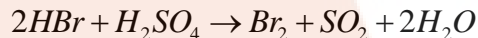
(a) In the manufacture of benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant because of the following reasons.

- In a neutral medium, OH^- ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.
- KMnO_4 and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol, KMnO_4 and toluene can react at a faster rate.

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

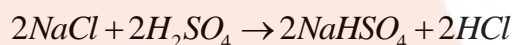


(b) When conc. H_2SO_4 is added to an inorganic mixture containing bromide, initially HBr is produced. HBr , being a strong reducing agent, reduces H_2SO_4 to SO_2 with the evolution of red vapour of bromine.

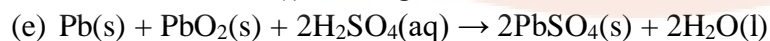
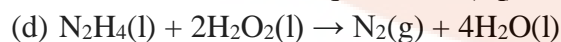
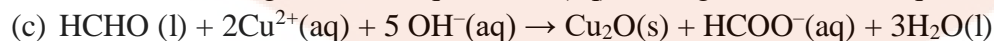
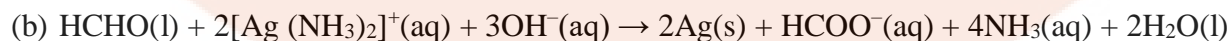
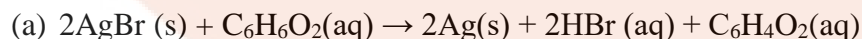


(red vapour)

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

**Question 13:**

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



Solution 13:

(a) Oxidised substance $\rightarrow C_6H_6O_2$

Reduced substance $\rightarrow AgBr$

Oxidising agent $\rightarrow AgBr$

Reducing agent $\rightarrow C_6H_6O_2$

(b) Oxidised substance $\rightarrow HCHO$

Reduced substance $\rightarrow [Ag(NH_3)_2]^+$

Oxidising agent $\rightarrow [Ag(NH_3)_2]^+$

Reducing agent $\rightarrow HCHO$

(c) Oxidised substance $\rightarrow HCHO$

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

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

Reducing agent $\rightarrow HCHO$

(d) Oxidised substance $\rightarrow N_2H_4$

Reduced substance $\rightarrow H_2O_2$

Oxidising agent $\rightarrow H_2O_2$

Reducing agent $\rightarrow N_2H_4$

(e) Oxidised substance $\rightarrow Pb$

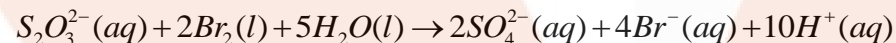
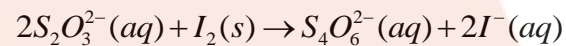
Reduced substance $\rightarrow PbO_2$

Oxidising agent $\rightarrow PbO_2$

Reducing agent $\rightarrow Pb$

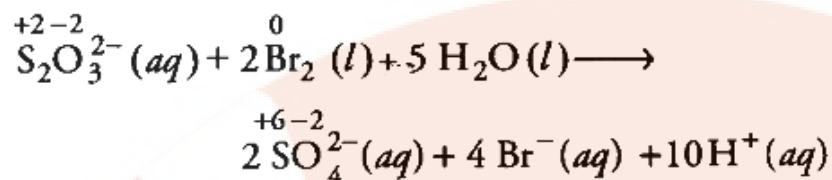
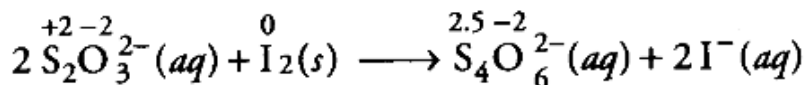
Question 14:

Consider the reactions:



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

Solution 14:



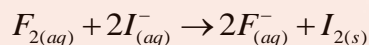
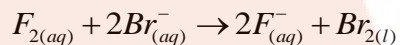
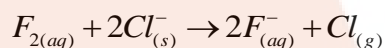
Bromine is a stronger oxidizing agent when compared to I₂. It oxidises the S of S₂O₃²⁻ to a higher oxidation state +6 in SO₄²⁻. While I₂ oxidises S of S₂O₃²⁻ to a lower oxidation state 2.5 in S₄O₆²⁻. That's why same reductant, thiosulphate react differently with bromine and iodine.

Question 15:

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

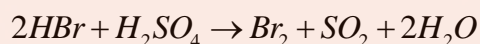
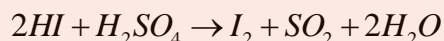
Solution 15:

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

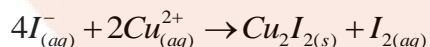


On the other hand, Cl₂, Br₂, and I₂ cannot oxidize F⁻ to F₂. The oxidizing power of halogens increases in the order of I₂ < Br₂ < Cl₂ < F₂. Hence, fluorine is the best oxidant among halogens.

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



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

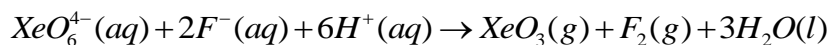


Hence, hydroiodic acid is the best reductant among hydrohalic compounds.

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

Question 16:

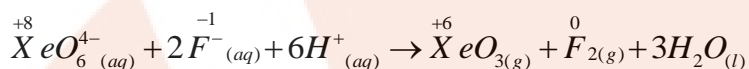
Why does the following reaction occur?



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

Solution 16:

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

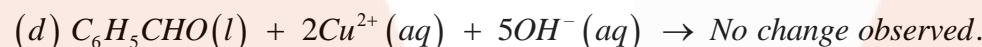
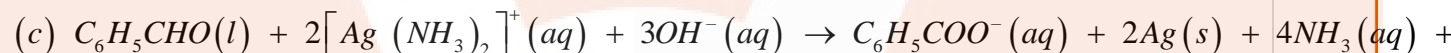
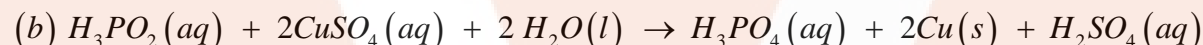
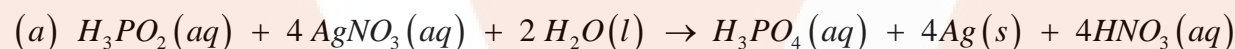


In this reaction, the oxidation number (O.N.) of Xe decreases from +8 in XeO_6^{4-} to +6 in XeO_3 and the O.N. of F increases from –1 in F^{-} to 0 in F_2 .

Hence, we can conclude that Na_4XeO_6 is a stronger oxidising agent than F^{-} .

Question 17:

Consider the reactions:



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

Solution 17:

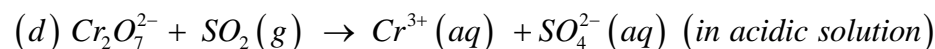
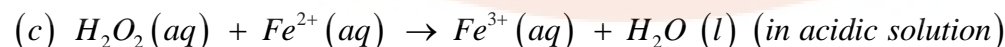
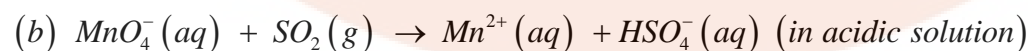
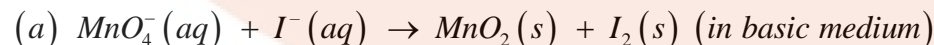
Ag^{+} and Cu^{2+} act as oxidising agents in reactions (a) and (b) respectively.

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

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

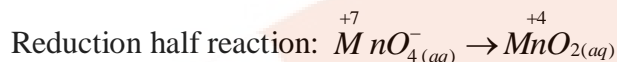
Question 18:

Balance the following redox reactions by ion-electron method:

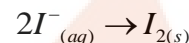


Solution 18:

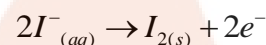
(a) Step 1: The two half reactions involved in the given reaction are:

**Step 2:**

Balancing I in the oxidation half reaction, we have:

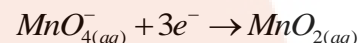


Now, to balance the charge, we add $2e^{-}$ to the RHS of the reaction.

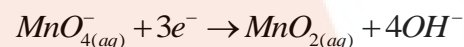
**Step 3:**

In the reduction half reaction, the oxidation state of Mn has reduced from +7 to +4.

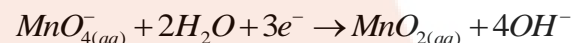
Thus, 3 electrons are added to the LHS of the reaction.



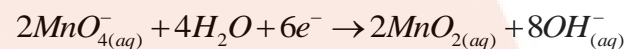
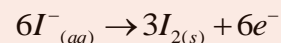
Now, to balance the charge, we add $4OH^{-}$ ions to the RHS of the reaction as the reaction is taking place in a basic medium.

**Step 4:**

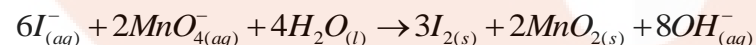
In this equation, there are 6 O atoms on the RHS and 4 O atoms on the LHS. Therefore, two water molecules are added to the LHS.

**Step 5:**

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

**Step 6:**

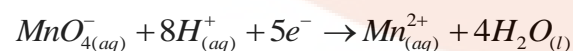
Adding the two half reactions, we have the net balanced redox reaction as:



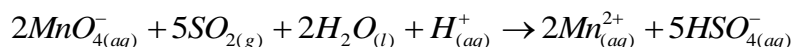
(b) Following the steps as in part (a), we have the oxidation half reaction as:



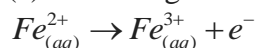
And the reduction half reaction as:



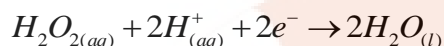
Multiplying the oxidation half reaction by 5 and the reduction half reaction by 2, and then by adding them, we have the net balanced redox reaction as:



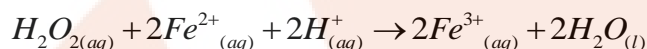
(c) Following the steps as in part (a), we have the oxidation half reaction as:



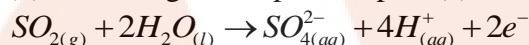
And the reduction half reaction as:



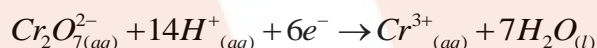
Multiplying the oxidation half reaction by 2 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:



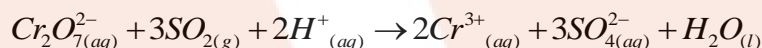
(d) Following the steps as in part (a), we have the oxidation half reaction as:



And the reduction half reaction as:

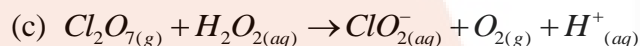
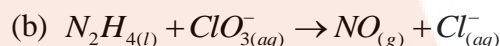
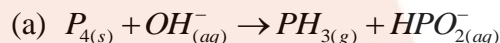


Multiplying the oxidation half reaction by 3 and then adding it to the reduction half reaction, we have the net balanced redox reaction as:



Question 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.

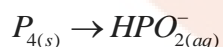


Solution 19:

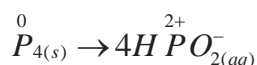
(a) The O.N. (oxidation number) of P decreases from 0 in P_4 to -3 in PH_3 and increases from 0 in P_4 to $+2$ in HPO_2^- . Hence, P_4 acts both as an oxidizing agent and a reducing agent in this reaction.

Ion-electron method:

The oxidation half equation is:



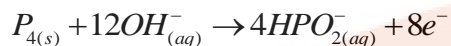
The P atom is balanced as:



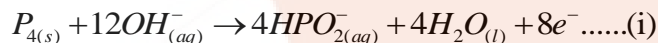
The O.N. is balanced by adding 8 electrons as:



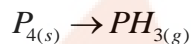
The charge is balanced by adding $12OH^-$ as:



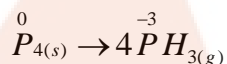
The H and O atoms are balanced by adding $4H_2O$ as:



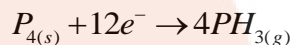
The reduction half equation is:



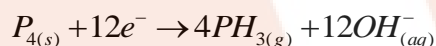
The P atom is balanced as



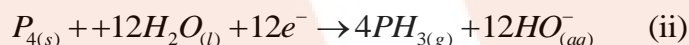
The O.N. is balanced by adding 12 electrons as:



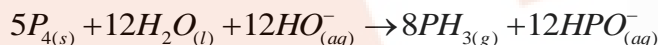
The charge is balanced by adding $12OH^-$ as:



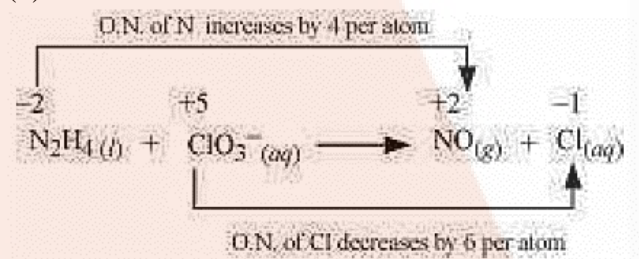
The O and H atoms are balanced by adding $12H_2O$ as:



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



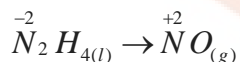
(b)



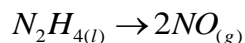
The oxidation number of N increases from -2 in N_2H_4 to $+2$ in NO and the oxidation number of Cl decreases from $+5$ in ClO_3^- to -1 in Cl^- . Hence, in this reaction, N_2H_4 is the reducing agent and ClO_3^- is the oxidizing agent.

Ion–electron method:

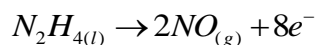
The oxidation half equation is:



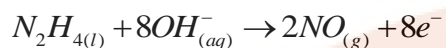
The N atoms are balanced as:



The oxidation number is balanced by adding 8 electrons as:



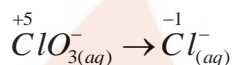
The charge is balanced by adding 8 OH⁻ ions as:



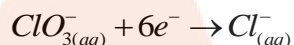
The O atoms are balanced by adding 6H₂O as:



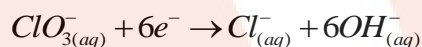
The reduction half equation is:



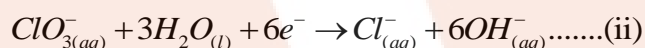
The oxidation number is balanced by adding 6 electrons as:



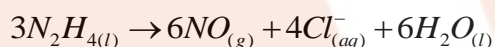
The charge is balanced by adding 6OH⁻ ions as:



The O atoms are balanced by adding 3H₂O as:



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

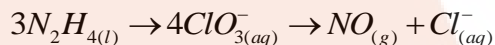


Oxidation number method:

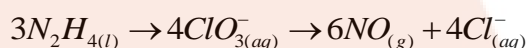
Total decrease in oxidation number of N = 2 × 4 = 8

Total increase in oxidation number of Cl = 1 × 6 = 6

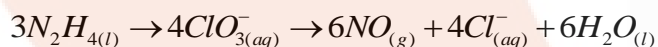
On multiplying N₂H₄ with 3 and ClO₃⁻ with 4 to balance the increase and decrease in O.N., we get:



The N and Cl atoms are balanced as:

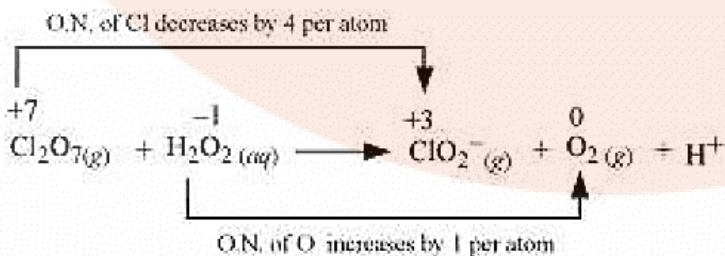


The O atoms are balanced by adding 6H₂O as:



This is the required balanced equation.

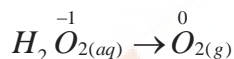
(c)



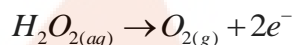
The oxidation number of Cl decreases from + 7 in Cl_2O_7 to + 3 in ClO_2^- and the oxidation number of O increases from – 1 in H_2O_2 to zero in O_2 . Hence, in this reaction, Cl_2O_7 is the oxidizing agent and H_2O_2 is the reducing agent.

Ion–electron method:

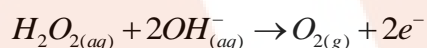
The oxidation half equation is:



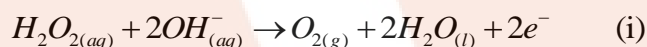
The oxidation number is balanced by adding 2 electrons as:



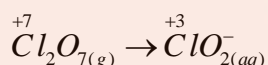
The charge is balanced by adding 2OH^- ions as:



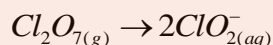
The oxygen atoms are balanced by adding $2\text{H}_2\text{O}$ as:



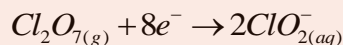
The reduction half equation is:



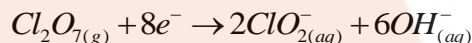
The Cl atoms are balanced as:



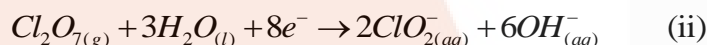
The oxidation number is balanced by adding 8 electrons as:



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



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



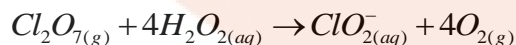
The balanced equation can be obtained by multiplying equation (i) with 4 and adding equation (ii) to it as: $\text{Cl}_2\text{O}_{7(g)} + 4\text{H}_2\text{O}_{2(aq)} + 2\text{OH}^-_{(aq)} \rightarrow 2\text{ClO}_{2(aq)}^- + 4\text{O}_{2(g)} + 5\text{H}_2\text{O}_{(l)}$

Oxidation number method:

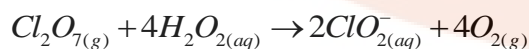
Total decrease in oxidation number of $\text{Cl}_2\text{O}_7 = 4 \times 2 = 8$

Total increase in oxidation number of $\text{H}_2\text{O}_2 = 2 \times 1 = 2$

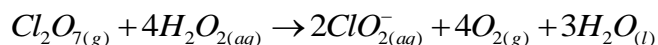
By multiplying H_2O_2 and O_2 with 4 to balance the increase and decrease in the oxidation number, we get:



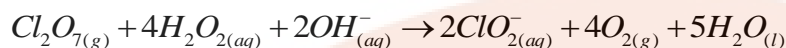
The Cl atoms are balanced as:



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



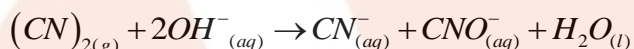
The H atoms are balanced by adding 2OH^- and $2\text{H}_2\text{O}$ as:



This is the required balanced equation.

Question 20:

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



Solution 20:

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

These are obtained as shown below:

Let the oxidation number of C be x .



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

$$\therefore x = 3$$



$$x - 3 = -1$$

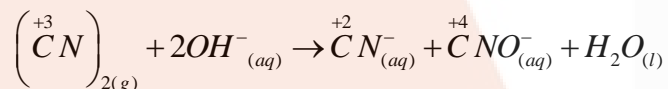
$$\therefore x = 2$$



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

$$\therefore x = 4$$

The oxidation number of carbon in the various species is:



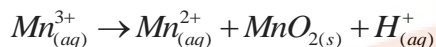
It can be easily observed that the same compound is being reduced and oxidized simultaneously in the given equation. Reactions in which the same compound is reduced and oxidised is known as disproportionation reactions. Thus, it can be said that the alkaline decomposition of cyanogen is an example of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

Question 21:

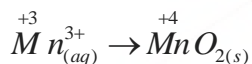
The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+} , MnO_2 , and H^+ ion. Write a balanced ionic equation for the reaction.

Solution 21:

The given reaction can be represented as:



The oxidation half equation is:



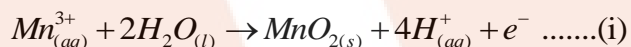
The oxidation number is balanced by adding one electron as:



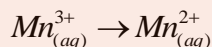
The charge is balanced by adding $4H^{+}$ ions as:



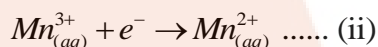
The O atoms and H^{+} ions are balanced by adding $2H_2O$ molecules as:



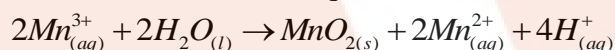
The reduction half equation is:



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 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.

Solution 22:

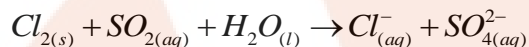
- F exhibits only negative oxidation state of -1 .
- Cs exhibits positive oxidation state of $+1$.
- I exhibit both positive and negative oxidation states. It exhibits oxidation states of -1 , $+1$, $+3$, $+5$, and $+7$.
- The oxidation state of Ne is zero. It exhibits neither negative nor positive oxidation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

Question 23:

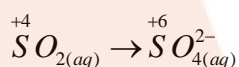
Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Solution 23:

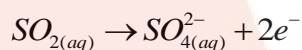
The given redox reaction can be represented as:



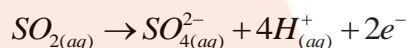
The oxidation half reaction is:



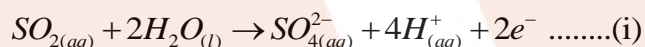
The oxidation number is balanced by adding two electrons as:



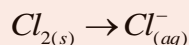
The charge is balanced by adding $4H^+$ ions as:



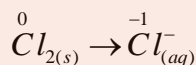
The O atoms and H^+ ions are balanced by adding $2H_2O$ molecules as:



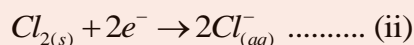
The reduction half reaction is:



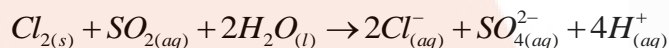
The chlorine atoms are balanced as:



The oxidation number is balanced by adding electrons



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

**Question 24:**

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

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

Solution 24:

In disproportionation reactions, one of the reacting substances always contains an element that

can exist in at least three oxidation states.

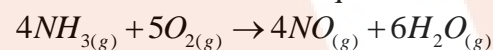
- (a) P, Cl, and S can show disproportionation reactions as these elements can exist in three or more oxidation states.
- (b) Mn, Cu, and Ga can show disproportionation reactions as these elements can exist in three or more oxidation states.

Question 25:

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

Solution 25:

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



$$\begin{array}{cccc} 4 \times 17g & 5 \times 32g & 4 \times 30g & 6 \times 18g \\ = 68g & = 160g & = 120g & = 108g \end{array}$$

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

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

But the available amount of O_2 is 20 g.

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

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

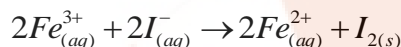
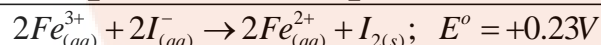
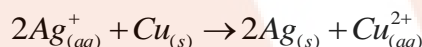
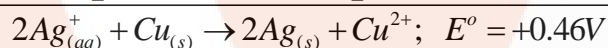
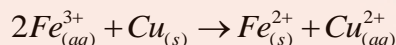
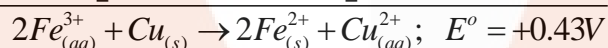
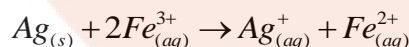
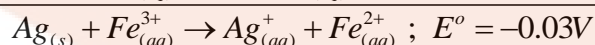
Therefore, 20 g of O_2 gives $\frac{120 \times 20}{160}$ g of N, or 15 g of NO.

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

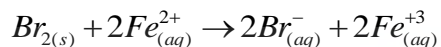
Question 26:

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

- (a) $Fe^{3+}(aq)$ and $I^-(aq)$
- (b) $Ag^+(aq)$ and $Cu(s)$
- (c) $Fe^{3+}(aq)$ and $Cu(s)$

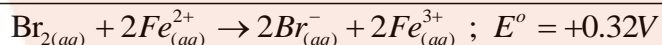
(d) Ag(s) and $\text{Fe}^{3+}(\text{aq})$ (e) $\text{Br}_2(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$ **Solution 26:**(a) The possible reaction between $\text{Fe}^{3+}_{(\text{aq})} + \text{I}^{-}_{(\text{aq})}$ is given by,Oxidation half equation: $2\text{I}^{-}_{(\text{aq})} \rightarrow \text{I}_{2(\text{s})} + 2\text{e}^{-}$; $E^{\circ} = -0.54\text{V}$ Reduction half equation: $[\text{Fe}^{3+}_{(\text{aq})} + \text{e}^{-} \rightarrow \text{Fe}^{2+}_{(\text{aq})}] \times 2$; $E^{\circ} = +0.77\text{V}$  E° for the overall reaction is positive. Thus, the reaction between $\text{Fe}^{3+}_{(\text{aq})}$ and $\text{I}^{-}_{(\text{aq})}$ is feasible.(b) The possible reaction between $\text{Ag}^{+}_{(\text{aq})} + \text{Cu}_{(\text{s})}$ is given by,Oxidation half equation: $\text{Cu}_{(\text{s})} \rightarrow \text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-}$; $E^{\circ} = -0.34\text{V}$ Reduction half equation: $[\text{Ag}^{+}_{(\text{aq})} + \text{e}^{-} \rightarrow \text{Ag}_{(\text{s})}] \times 2$; $E^{\circ} = +0.80\text{V}$  E° positive for the overall reaction is positive. Hence, the reaction between $\text{Ag}^{+}_{(\text{aq})}$ and $\text{Cu}_{(\text{s})}$ is feasible.(c) The possible reaction between $\text{Fe}^{3+}_{(\text{aq})}$ and $\text{Cu}_{(\text{s})}$ is given by,Oxidation half equation: $\text{Cu}_{(\text{s})} \rightarrow \text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-}$; $E^{\circ} = -0.34\text{V}$ Reduction half equation: $[\text{Fe}^{3+}_{(\text{aq})} + \text{e}^{-} \rightarrow \text{Fe}^{2+}_{(\text{aq})}] \times 2$; $E^{\circ} = +0.77\text{V}$  E° positive for the overall reaction is positive. Hence, the reaction between $\text{Fe}^{3+}_{(\text{aq})}$ and $\text{Cu}_{(\text{s})}$ is feasible.(d) The possible reaction between $\text{Ag}_{(\text{s})}$ and $\text{Fe}^{3+}_{(\text{aq})}$ is given by,Oxidation half equation: $\text{Ag}_{(\text{s})} \rightarrow \text{Ag}^{+}_{(\text{aq})} + \text{e}^{-}$; $E^{\circ} = -0.80\text{V}$ Reduction half equation: $\text{Fe}^{3+}_{(\text{aq})} + \text{e}^{-} \rightarrow \text{Fe}^{2+}_{(\text{aq})}$; $E^{\circ} = +0.77\text{V}$ Here, E° for the overall reaction is negative. Hence, the reaction between $\text{Ag}_{(\text{s})}$ and $\text{Fe}^{3+}_{(\text{aq})}$ is not feasible.

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



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

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



Here, E° for the overall reaction is positive. Hence, the reaction between $Br_{2(aq)}$ and $Fe_{(aq)}^{2+}$ is feasible.

Question 27:

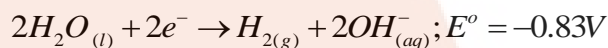
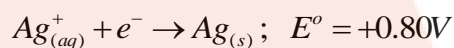
Predict the products of electrolysis in each of the following:

- (i) An aqueous solution of $AgNO_3$ with silver electrodes
- (ii) An aqueous solution $AgNO_3$ with platinum electrodes
- (iii) A dilute solution of H_2SO_4 with platinum electrodes
- (iv) An aqueous solution of $CuCl_2$ with platinum electrodes.

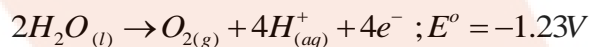
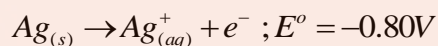
Solution 27:

(i) $AgNO_3$ ionizes in aqueous solutions to form Ag^{+} and NO_3^{-} ions.

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



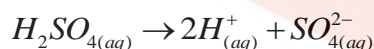
Hence, Ag^{+} ions are reduced at the cathode. Similarly, Ag metal or H_2O molecules can be oxidized at the anode. But the oxidation potential of Ag is higher than that of H_2O molecules.



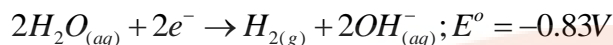
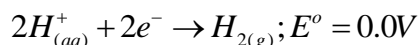
Therefore, Ag metal gets oxidized at the anode.

(ii) Pt cannot be oxidized easily. Hence, at the anode, oxidation of water occurs to liberate O_2 . At the cathode, Ag^{+} ions are reduced and get deposited.

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



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.

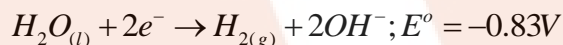
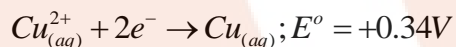
But the oxidation of SO_4^{2-} involves breaking of more bonds than that of H_2O molecules.

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:

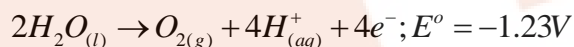
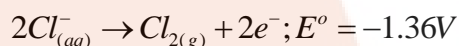


On electrolysis, either of Cu^{2+} ions or H_2O molecules can get reduced at the cathode. But the reduction potential of Cu^{2+} is more than that of H_2O molecules.



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.

Question 28:

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

Al, Cu, Fe, Mg and Zn.

Solution 28:

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

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

Hence, we can say that Mg can displace Al from its salt solution, but Al cannot displace Mg.

Thus, the order in which the given metals displace each other from the solution of their salts is given below:

$Mg > Al > Zn > Fe, > Cu$

Question 29:

Given the standard electrode potentials,

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

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

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

Arrange these metals in their increasing order of reducing power.

Solution 29:

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

Question 30:

Depict the galvanic cell in which the reaction

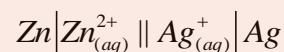


) takes place, further show:

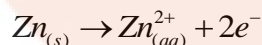
- (i) which of the electrode is negatively charged,
- (ii) the carriers of the current in the cell, and
- (iii) individual reaction at each electrode.

Solution 30:

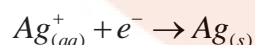
The galvanic cell corresponding to the given redox reaction can be represented as:



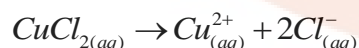
- (i) Zn electrode is negatively charged because at this electrode, Zn oxidizes to Zn^{2+} and the leaving electrons accumulate on this electrode.
- (ii) Ions are the carriers of current in the cell.
- (iii) The reaction taking place at Zn electrode can be represented as:



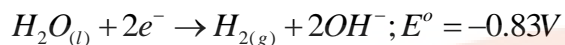
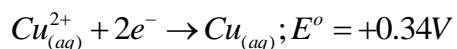
And the reaction taking place at Ag electrode can be represented as:



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

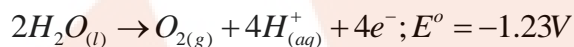
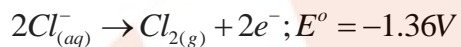


On electrolysis, either of Cu^{2+} ions or H_2O molecules can get reduced at the cathode. But the reduction potential of Cu^{2+} is more than that of H_2O molecules.



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.

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