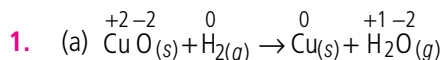


Redox Reactions

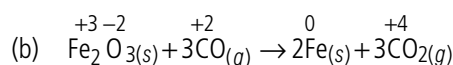


ANSWERS

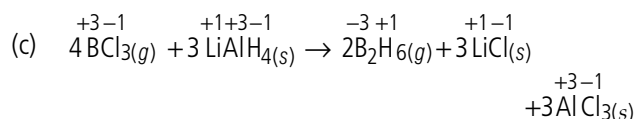
TOPIC 1



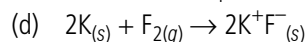
O.N. of Cu decreases from +2 in CuO to 0 in Cu but that of H increases from 0 in H₂ to +1 in H₂O. Therefore, CuO is reduced to Cu but H₂ is oxidised to H₂O. Thus, this is a redox reaction.



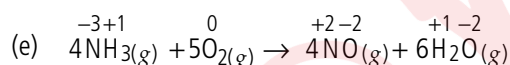
O.N. of Fe decreases from +3 in Fe₂O₃ to 0 in Fe while that of C increases from +2 in CO to +4 in CO₂. Therefore, Fe₂O₃ is reduced and CO is oxidised. Thus, this is a redox reaction.



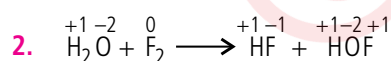
O.N. of B decreases from +3 in BCl₃ to -3 in B₂H₆ while that of H increases from -1 in LiAlH₄ to +1 in B₂H₆. Therefore, BCl₃ is reduced while LiAlH₄ is oxidised. Thus, this is a redox reaction.



Each K atom has lost one electron to form K⁺ while F₂ has gained two electrons to form two F⁻ ions. Therefore, K is oxidised while F₂ is reduced. Thus, it is a redox reaction.

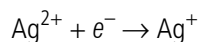


O.N. of N increases from -3 in NH₃ to +2 in NO while that of O decreases from 0 in O₂ to -2 in NO or H₂O. Therefore, NH₃ is oxidised while O₂ is reduced. Thus, it is a redox reaction.



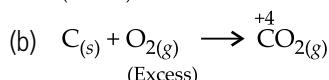
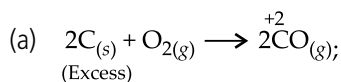
Since fluorine can undergo oxidation as well as reduction hence, it is an example of redox reaction.

3. In AgF₂ oxidation state of Ag is +2 which is very unstable. Therefore, it quickly, accepts an electron to form the more stable +1 oxidation state.

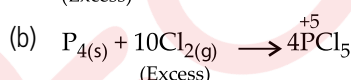
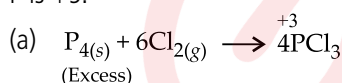


Therefore, AgF₂, if formed, will act as a strong oxidising agent.

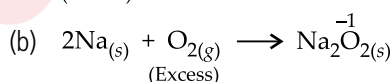
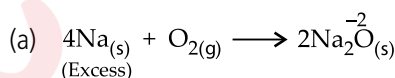
4. (i) C is a reducing agent while O₂ is an oxidising agent. If excess of carbon is burnt in a limited supply of O₂, CO is formed in which oxidation state of C is +2 but when O₂ is in excess CO formed gets oxidised to CO₂ in which oxidation state of C is +4



(ii) P₄ is a reducing agent while Cl₂ is an oxidising agent. When excess of P₄ is used, PCl₃ is formed in which the oxidation state of P is +3. When excess of Cl₂ is used, the initially formed PCl₃ reacts further to form PCl₅ in which the oxidation state of P is +5.

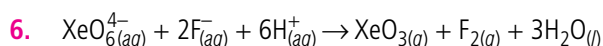


(iii) Na is a reducing agent and O₂ is an oxidising agent. When excess of Na is used, sodium oxide is formed in which the oxidation state of O is -2. If, excess of O₂ is used, Na₂O₂ is formed in which the oxidation state of O is -1.



5.

	Reducing agent (Substance oxidised)	Oxidising agent (Substance reduced)
(a)	C ₆ H ₆ O _{2(aq)}	AgBr _(s)
(b)	HCHO _(l)	[Ag(NH ₃) ₂] ⁺ _(aq)
(c)	HCHO _(l)	Cu ²⁺ _(aq)
(d)	N ₂ H _{4(l)}	H ₂ O _{2(l)}
(e)	Pb _(s)	PbO _{2(s)}



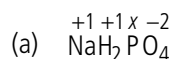
In this reaction, O.N. of Xe decreases from +8 in XeO₆⁴⁻ to +6 in XeO₃ while that of F increases from -1 in F⁻ to 0 in F₂. Therefore, XeO₆⁴⁻ is reduced while F⁻ is oxidised. From this reaction it is concluded that Na₄XeO₆ is a stronger oxidising agent than F₂.

7. Reactions (a) and (b) indicate that both AgNO₃ and CuSO₄ oxidise H₃PO₂ to H₃PO₄. Hence, both are oxidising agents.

Reactions (c) and (d) suggest that $[\text{Ag}(\text{NH}_3)_2]^+$ oxidises $\text{C}_6\text{H}_5\text{CHO}$ (benzaldehyde) to $\text{C}_6\text{H}_5\text{COO}^-$ (benzoate ion) but Cu^{2+} ions cannot oxidise $\text{C}_6\text{H}_5\text{CHO}$ to $\text{C}_6\text{H}_5\text{COO}^-$. Therefore, from the above reactions, we infer that Ag^+ ion is a stronger oxidising agent than Cu^{2+} ion.

TOPIC 2

1. Let the oxidation no. of underlined element in all the given compounds = x

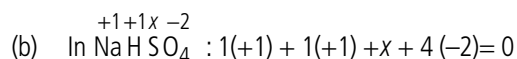


Since the sum of oxidation number of various atoms in NaH_2PO_4 (neutral) is zero.

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

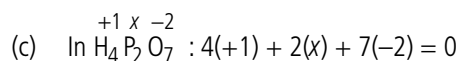
$$x - 5 = 0 \quad \text{or} \quad x = +5$$

Thus, the oxidation number of P in $\text{NaH}_2\text{PO}_4 = +5$



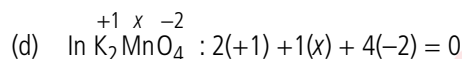
$$\text{or} \quad x = +6$$

Thus, the oxidation number of S in $\text{NaHSO}_4 = +6$



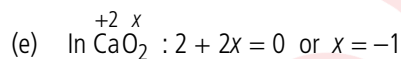
$$\text{or} \quad x = +5$$

Thus, the oxidation number of P in $\text{H}_4\text{P}_2\text{O}_7 = +5$



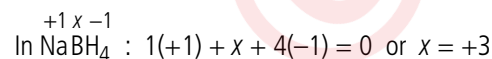
$$\text{or} \quad x = +6$$

Thus, the oxidation number of Mn in $\text{K}_2\text{MnO}_4 = +6$

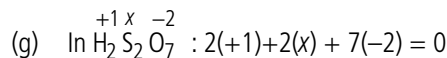


Thus, the oxidation number of oxygen in $\text{CaO}_2 = -1$

(f) In NaBH_4 , hydrogen is present as hydride ion. Therefore, its oxidation number is -1 . Thus,

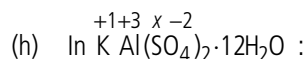


Thus, the oxidation number of B in $\text{NaBH}_4 = +3$



$$\text{or} \quad x = +6$$

Thus, the oxidation number of S in $\text{H}_2\text{S}_2\text{O}_7 = +6$



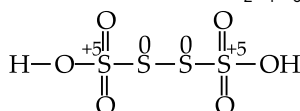
$$+1 + 3 + 2x + 8(-2) + 12 \times 0 = 0 \quad \text{or} \quad x = +6$$

Thus, the oxidation number of S in $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O} = +6$

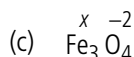
2. (a) In KI_3 , since the oxidation number of K is $+1$, therefore, the average oxidation number of iodine $= -1/3$. In the structure, $\text{K}^+[\text{I}-\text{I}-\text{I}]^-$, a coordinate bond is formed between I_2 molecule and I^- ion. The oxidation number of two iodine atoms forming

the I_2 molecule is zero while that of iodine ion forming the coordinate bond is -1 . Thus, the O.N. of three iodine atoms in KI_3 are $0, 0$ and -1 respectively.

(b) The structure of $\text{H}_2\text{S}_4\text{O}_6$ is shown below :



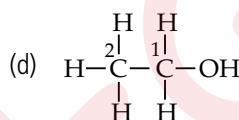
The O.N. of each of the S atoms linked with each other in the middle is zero while that of each of the remaining two S-atoms is $+5$.



Let O.N. of Fe = x , then $3x + 4(-2) = 0$

$$\text{or} \quad x = +\frac{8}{3} \quad (\text{average})$$

By stoichiometry Fe_3O_4 is $\text{Fe}^{+2}\text{O} \cdot \text{Fe}_2^{+3}\text{O}_3$. Thus, Fe has O.N. of $+2$ and $+3$.



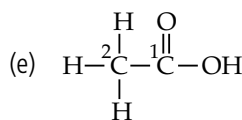
In this molecule, C-2 is attached to three H-atoms (less electronegative than carbon) and one CH_2OH group (more electronegativity than carbon).

Therefore O.N. of C-2 = $3(+1) + x + 1(-1) = 0$

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

C-1 is however, attached to one OH (charge = -1) and one CH_3 (charge = $+1$), and two H-atoms, O.N. of $+1$.

Therefore, O.N. of C-1 = $1(+1) + 2(+1) + x + 1(-1) = 0$ or $x = -2$



In this molecule, C-2 is attached to three H-atoms (less electronegative than carbon) and one $-\text{COOH}$ group (more electronegativity than carbon).

Therefore, O.N. of C-2 = $3(+1) + x + 1(-1) = 0$

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

C-1 is, however, attached to one oxygen atom by a double bond, one OH (charge = -1) and one CH_3 (charge = $+1$) group, therefore, O.N. of C-1 = $1(+1) + x + 1(-2) + 1(-1) = 0$

$$\text{or} \quad x = +2$$

3. Oxidation number of sulphur in H_2SO_5 :

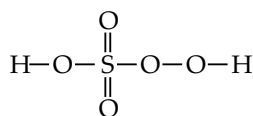
Let the oxidation number of S = x

$$\text{then } (+1) \times 2 + x + (-2) \times 5 = 0 \quad \text{or } 2 + x - 10 = 0$$

$$\Rightarrow x - 8 = 0$$

$$\therefore x = +8$$

The maximum O.N. of S cannot be more than 6 since it has only 6 electrons in the valence shell. This fallacy is overcome if we calculate the O.N. of sulphur by chemical bonding method. The structure of H_2SO_5 is



It has two peroxide oxygen with O.N. = -1 and three oxygens with O.N. = -2

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

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

Thus, O.N. of sulphur in $\text{H}_2\text{SO}_5 = +6$

Oxidation number of chromium in $\text{Cr}_2\text{O}_7^{2-}$:

Let the oxidation number of chromium = x

$$\therefore 2x + 7(-2) = -2 \Rightarrow 2x - 14 = -2$$

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

$$\Rightarrow 2x = +12 \Rightarrow x = +6$$

Thus, the oxidation number of chromium = +6

Oxidation number of nitrogen in NO_3^- :

Let the oxidation number of nitrogen = x

$$\text{then } x + (-2) \times 3 = -1 \Rightarrow x - 6 = -1$$

$$\Rightarrow x = -1 + 6 \Rightarrow x = +5$$

Thus, the oxidation number of nitrogen = +5

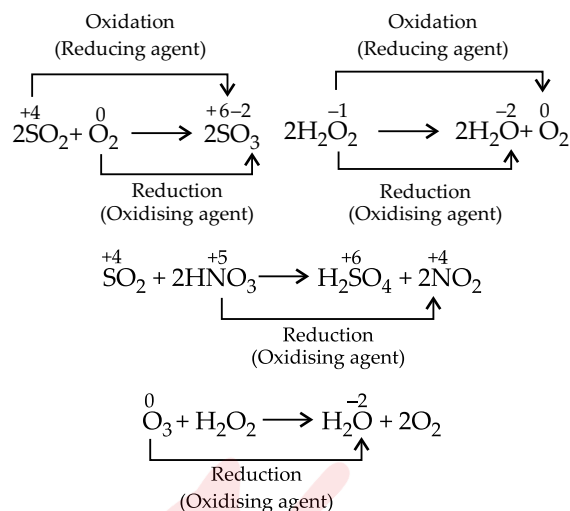
4. (a) HgCl_2 (b) NiSO_4
 (c) SnO_2 (d) Ti_2SO_4
 (e) $\text{Fe}_2(\text{SO}_4)_3$ (f) Cr_2O_3

5.

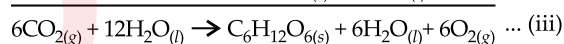
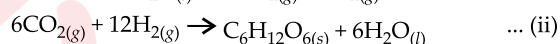
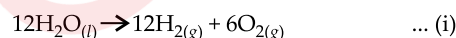
Compound	O.N. of Carbon	Compound	O.N. of Nitrogen
CH_4	-4	$\text{NH}_3, \text{NH}_4\text{Cl}$	-3
$\text{CH}_3 - \text{CH}_3$	-3	$\text{NH}_2 - \text{NH}_2$	-2
$\text{CH}_2 = \text{CH}_2$ or CH_3Cl	-2	$\text{NH} = \text{NH},$ NH_2OH	-1
$\text{CH} \equiv \text{CH}$	-1	$\text{N} \equiv \text{N}$	0
$\text{C}_6\text{H}_{12}\text{O}_6$	0	N_2O	+1
C_6Cl_6	+1	NO	+2
CHCl_3	+2	$\text{N}_2\text{O}_3, \text{HNO}_2$	+3
$(\text{COOH})_2$	+3	N_2O_4	+4
$\text{CCl}_4, \text{CO}_2$	+4	$\text{N}_2\text{O}_5, \text{HNO}_3$	+5

6. The oxidation state of sulphur in sulphur dioxide is +4. It can be oxidised to +6 oxidation state or reduced to +2. Therefore, sulphur dioxide acts as a reducing agent as well as oxidising agent. Similarly, the oxidation state of oxygen in hydrogen peroxide is -1. It can be oxidised to O_2 (zero oxidation state) or reduced to H_2O or OH^- (-2 oxidation state) and therefore, acts as reducing as well as oxidising agents.

However, both ozone and nitric acid can only decrease their oxidation number and therefore, act only as oxidising agents.

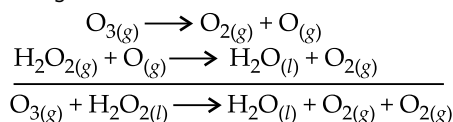


7. (a) Most accepted mechanism of photosynthesis is supposed to occur in two steps. In the first step, H_2O decomposes to give H_2 and O_2 in presence of chlorophyll and the H_2 thus produced reduces CO_2 to $\text{C}_6\text{H}_{12}\text{O}_6$ and H_2O molecules are also produced as shown below:



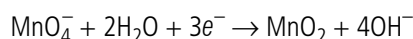
Therefore, it is more appropriate to write the equation for photosynthesis as (iii) because it emphasises that 12 molecules of H_2O are used per molecule of carbohydrate formed and 6 molecules of H_2O are produced during the process.

(b) The purpose of writing O_2 two times suggests that O_2 is being obtained from each of the two reactants.

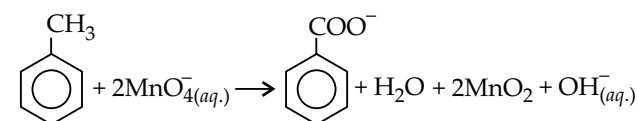


The path of reactions (a) and (b) can be determined by tracer technique method using labelled H or O, H_2O^{18} or D_2O in reaction (a) or by using $\text{H}_2\text{O}_2^{18}$ or O_3^{18} in reaction (b).

8. (a) In neutral medium, KMnO_4 acts as an oxidant as follows:



In laboratory, alkaline KMnO_4 is used to oxidise toluene to benzoic acid.



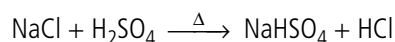
Toluene

In industry alcoholic KMnO_4 is preferred due to following reasons:

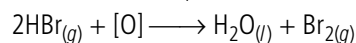
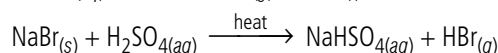
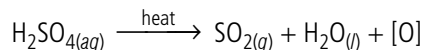
1. Cost of adding a base or an acid is avoided as a base (OH^- ions) is produced during the reaction itself.

2. An organic polar solvent, ethyl alcohol, helps in mixing of two reactants KMnO_4 (due to its polar nature) and toluene (because, an organic compound).

(b) When a chloride such as NaCl is heated with conc. H_2SO_4 , HCl is evolved which is not a good reducing agent and is not oxidised with conc. H_2SO_4 .

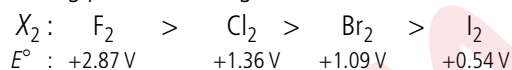


But when NaBr is heated with conc. H_2SO_4 , HBr produced being a good reducing agent is oxidised by conc. H_2SO_4 to give Br_2 (red vapours).

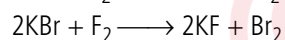
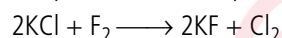


9. The average O.N. of S in $\text{S}_2\text{O}_3^{2-}$ is +2 while in $\text{S}_4\text{O}_6^{2-}$ it is +2.5. The O.N. of S in SO_4^{2-} is +6. Since Br_2 is a stronger oxidising agent, it oxidises S of $\text{S}_2\text{O}_3^{2-}$ to a higher oxidation state of +6 and hence forms SO_4^{2-} ion. I_2 , however, being a weaker oxidising agent oxidises S of $\text{S}_2\text{O}_3^{2-}$ ion to a lower oxidation state of +2.5 in $\text{S}_4\text{O}_6^{2-}$ ion.

10. The halogens (X_2) have strong electron accepting tendency and have positive standard oxidation potential values. They are therefore, powerful oxidising agents. The decreasing order of oxidising powers of halogens is :



Fluorine is the strongest oxidising agent (oxidant) because it can liberate the other halogens from their respective compounds. For example,



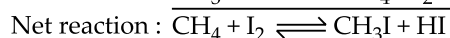
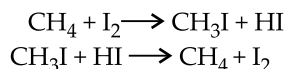
Among halogen acids (HX), the HI is the strongest reducing agent or reductant because it has minimum bond dissociation energy.

Halogen acid : HF HCl HBr HI

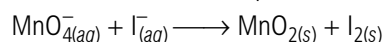
Bond dissociation : 566 431 366 299

enthalpy (kJ mol^{-1})

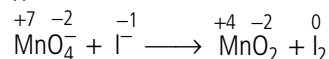
The iodination of methane is of reversible nature because HI produced in the reaction being a reducing agent converts iodomethane back to methane.



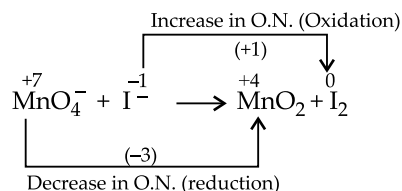
11. (a) The skeleton equation is :



(i) The O.N. of the atoms involved in the equation is :



(ii) The species involved in the oxidation and reduction half reaction :



(iii) Oxidation half reaction : $\text{I}^- \longrightarrow \text{I}_2$

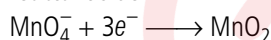
Reduction half reaction : $\text{MnO}_4^- \longrightarrow \text{MnO}_2$

(iv) Balancing the oxidation half reaction

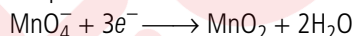


(v) Balancing the reduction half reaction

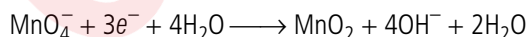
(1) As the decrease in O.N. is 3, therefore, adding 3e^- on the reactant side



(2) To balance the oxygen atoms, add two H_2O molecules on the product side



(3) To balance the charges, add 4 OH^- on the product side. Then to balance H atoms, add four H_2O molecules on the reactant side.

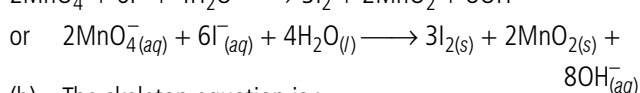
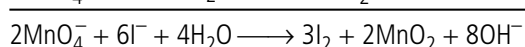
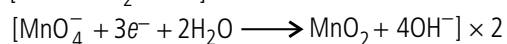
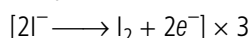


Thus, the reduction half reaction is balanced.

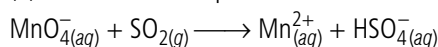
(vi) Adding the two half reactions

In order to equate the electrons, multiply eqn. (i) by 3 and eqn.

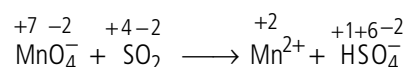
(ii) by 2. Add the two equations.



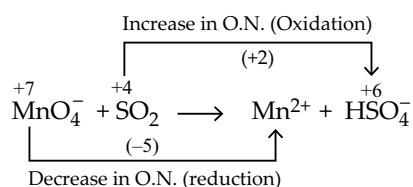
(b) The skeleton equation is :



(i) The O.N. of atoms involved in the equation is :



(ii) The species involved in the oxidation and reduction half reactions :



(iii) Oxidation half reaction : $\text{SO}_2 \longrightarrow \text{HSO}_4^-$

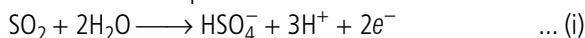
Reduction half reaction : $\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$

(iv) Balancing the oxidation half reaction

(1) As the increase in O.N. is 2, therefore, add two electrons on the product side to balance change in O.N.



(2) In order to balance the number of oxygen atoms, add two H_2O molecules on the reactant side and then to balance H atoms add 3H^+ on the product side.

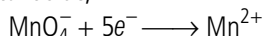


(v) Balancing the reduction half reaction

The reduction half reaction is :



(1) As the decrease in O.N. is 5, therefore add $5e^-$ on the reactant side,

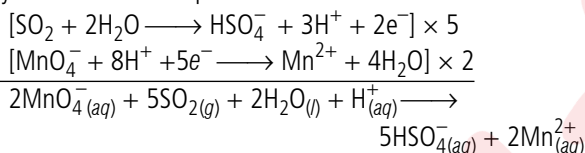


(2) In order to balance the no. of oxygen atoms, add four H_2O molecules on the product side and then to balance H atoms add 8H^+ on the reactant side.



(vi) Adding the two half, reactions

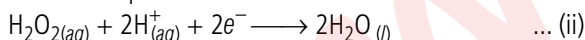
In order to equate the electrons, multiply eqn. (i) by 5 and eqn. (ii) by 2. Add the two eqns.



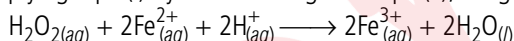
(c) Oxidation half equation :



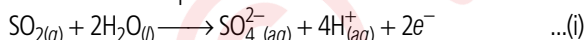
Reduction half equation :



Multiplying eqn. (i) by 2 and adding it to eqn. (ii), we get



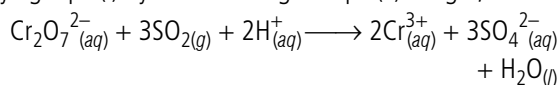
(d) Oxidation half equation :



Reduction half equation :

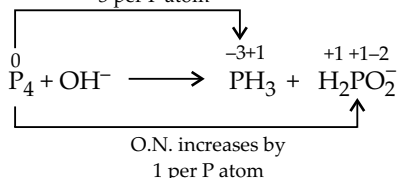


Multiplying eqn. (i) by 3 and adding to eqn. (ii) we get,



12. Oxidation number method :

(a) O.N. decreases by 3 per P atom

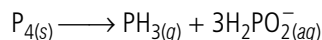


Total increase in O.N. of P from P_4 to $\text{H}_2\text{PO}_2^- = 1 \times 4 = 4$

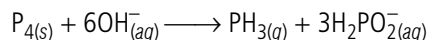
Total decrease in O.N. of P from P_4 to $\text{PH}_3 = 3 \times 4 = 12$

Therefore, to balance increase/decrease in O.N. multiply PH_3

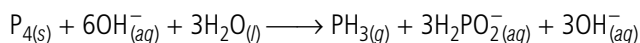
by 1 and H_2PO_2^- by 3, we have



To balance O atoms, add 6OH^- on the left side



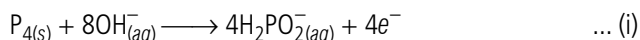
To balance H atoms, add $3\text{H}_2\text{O}$ to L.H.S. and 3OH^- to the R.H.S.



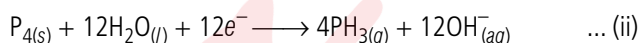
or, $\text{P}_{4(\text{s})} + 3\text{OH}^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l}) \longrightarrow \text{PH}_{3(\text{g})} + 3\text{H}_2\text{PO}_2^-(\text{aq})$

Ion electron method : The two half reactions are :

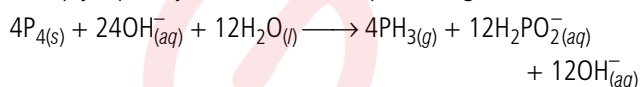
Oxidation half reaction :



Reduction half reaction :

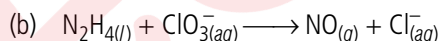


Multiply eq. (i) by 3 and add it to eq. (ii), we get

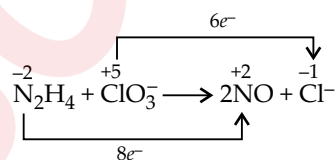


or, $\text{P}_{4(\text{s})} + 3\text{OH}^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l}) \longrightarrow \text{PH}_{3(\text{g})} + 3\text{H}_2\text{PO}_2^-(\text{aq})$

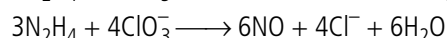
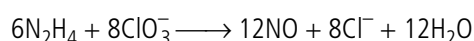
Reductant - phosphorus; oxidant-phosphorus



Oxidation number method :

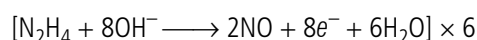


Net reaction is

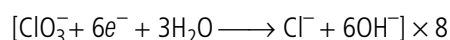


Ion-electron method :

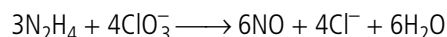
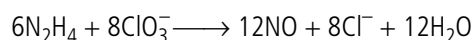
Oxidation half-reaction :



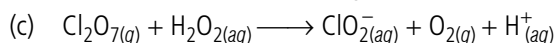
Reduction half-reaction :



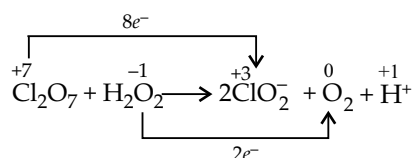
Net reaction is



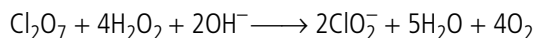
Reductant : N_2H_4 ; Oxidant : ClO_3^-



Oxidation number method :

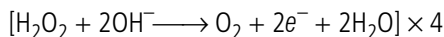


Net reaction is

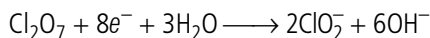


Ion-electron method :

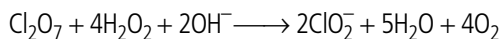
Oxidation half-reaction :



Reduction half-reaction :



Net reaction is

Reductant : H_2O_2 ; Oxidant : Cl_2O_7 **13.** Let x be the oxidation number of C.

$$\therefore \text{O.N. of C in cyanogen, } (\text{CN})_2 = 2(x - 3) = 0$$

$$\text{or } x = +3$$

$$\text{O.N. of C in cyanide ion, } \text{CN}^- = x - 3 = -1$$

$$\text{or } x = +2$$

$$\text{O.N. of C in cyanate ion, } \text{CNO}^- = x - 3 - 2 = -1$$

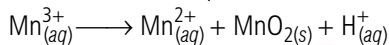
$$\text{or } x = +4$$

The four informations about the reaction are :

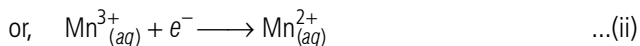
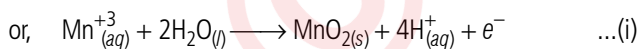
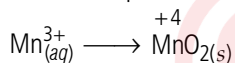
(i) The reaction involves decomposition of cyanogen, $(\text{CN})_2$ in the alkaline medium to cyanide ion, CN^- and cyanate ion, CNO^- .(ii) The O.N. of C decreases from +3 in $(\text{CN})_2$ to +2 in CN^- ion and increases from +3 in $(\text{CN})_2$ to +4 in CNO^- ion. Thus, cyanogen is simultaneously reduced to cyanide ion and oxidised to cyanate ion.

(iii) It is an example of disproportionation redox reaction.

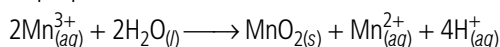
(iv) Cyanogen is a pseudohalogen while cyanide ion is a pseudohalide ion.

14. The skeletal equation is :

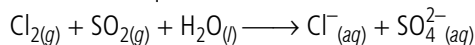
Oxidation half equation :



Adding Eq. (i) and Eq. (ii), the balanced equation for the disproportionation reaction is

**15.** (a) F : Fluorine is the most electronegative element and shows only -1 oxidation state.(b) Cs : Because of the presence of single electron in the valence shell, (alkali metals) Cs exhibits an oxidation state of $+1$ only.(c) I : Because of the presence of seven electrons in the valence shell, I shows an oxidation state of -1 in compounds with more electropositive elements (such as H, Na, K, Ca, etc.)and oxidation states of $+3$, $+5$, $+7$ in compounds of I with more electronegative elements (such as, O, F, etc.)

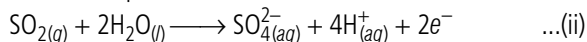
(d) Ne : It is an inert gas (with high ionization enthalpy and highly positive electron gain enthalpy) and hence, it exhibits neither negative nor positive oxidation states.

16. The skeletal equation is :

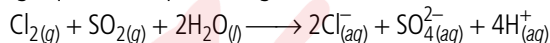
Reduction half equation :



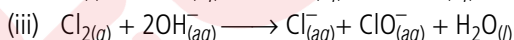
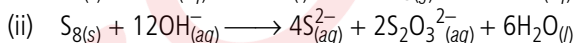
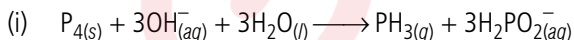
Oxidation half equation :



Adding Eq. (i) and Eq. (ii), we get,

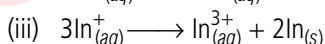
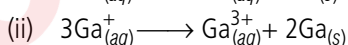
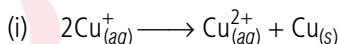
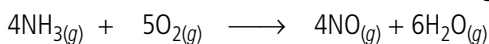


This represents the balanced redox reaction.

17. (a) The non-metals which show disproportionation reaction are N, P, S, Cl, Br and I.

Bromine and iodine give similar reactions as chlorine.

(b) The metals which undergo disproportionation reaction are :

Cu⁺, Ga⁺, In⁺, etc.**18.** The reaction involved in the manufacturing process is :

4 moles 5 moles 4 moles

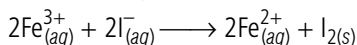
4 × 17 = 68 g 5 × 32 = 160 g 4 × 30 = 120 g

∴ 68 g of NH_3 will react with 160 g of O_2 to produce 120 g of $\text{NO}_{(g)}$

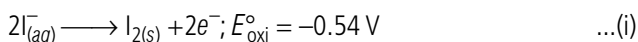
But oxygen which is actually available (20.0 g) is less than the amount which is required. Therefore, oxygen is the limiting reactant.

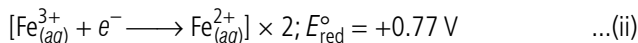
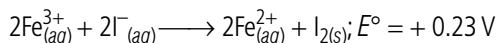
Now 160 g of O_2 will form $\text{NO} = 120$ g

$$\therefore 20 \text{ g of } \text{O}_2 \text{ will form } \text{NO} = \frac{(120)}{(160)} \times (20) = 15 \text{ g}$$

TOPIC 3**1.** (a) The possible reaction between $\text{Fe}^{3+}_{(aq)}$ and $\text{I}^-_{(aq)}$ is

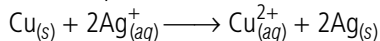
The above redox reaction can be split into the following two half reactions,

Oxidation :

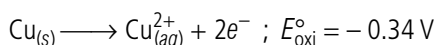
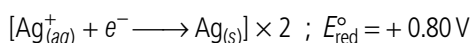
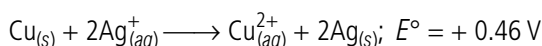
Reduction :**Overall reaction :**

Since the EMF for the above reaction is positive, therefore, the above reaction is feasible.

(b) The possible reaction between $\text{Ag}_{(s)}^{+}$ and $\text{Cu}_{(s)}$ is

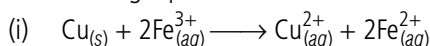


The above redox reaction can be split into the following two half reactions,

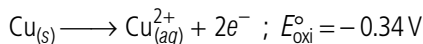
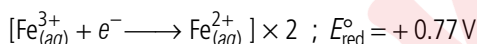
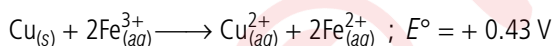
Oxidation :**Reduction :****Overall reaction :**

Since the EMF for the above reaction comes out to be positive, therefore, the above reaction is feasible.

(c) There are two probabilities for reaction between Cu and Fe^{3+} . The reaction between $\text{Fe}_{(aq)}^{3+}$ and $\text{Cu}_{(s)}$ occurs according to the following equation :

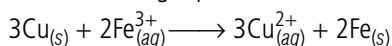


The above reaction can be split into the following two half reactions,

Oxidation :**Reduction :****Overall reaction :**

Since the EMF for the above reaction is positive, therefore, the above reaction is feasible.

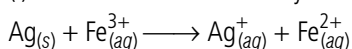
(ii) If the reaction between $\text{Fe}_{(aq)}^{3+}$ and $\text{Cu}_{(s)}$ occurs according to the following equation,



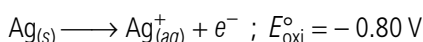
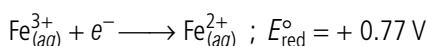
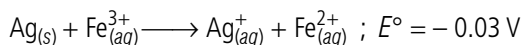
The EMF of the reaction comes out to be negative, *i.e.*, $-0.376 \text{ V} (-0.34 \text{ V} - 0.036 \text{ V})$ and hence, this reaction is not feasible.

(d) There are two possibilities for reaction between Ag and Fe^{3+} ion.

(i) If the reaction occurs by following equation,

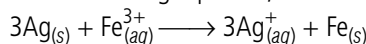


The reaction can be split into the following two half reactions,

Oxidation :**Reduction :****Overall reaction :**

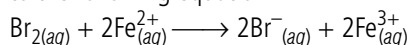
Since the EMF for the above reaction is negative, therefore, the above reaction is not feasible.

(ii) The reaction between $\text{Ag}_{(s)}$ and $\text{Fe}_{(aq)}^{3+}$ may occur according to the following equation,

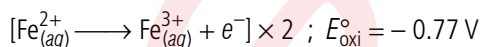
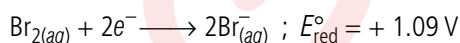
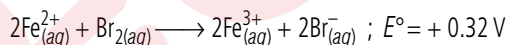


EMF of this reaction comes to be even more negative *i.e.*, -0.836 V , and hence this redox reaction is also not feasible.

(e) The reaction between $\text{Br}_{2(aq)}$ and $\text{Fe}_{(aq)}^{2+}$ occurs according to the following equation :

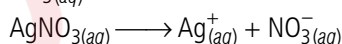


The above reaction can be split into the following two half reactions.

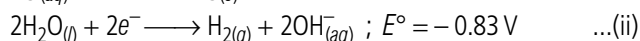
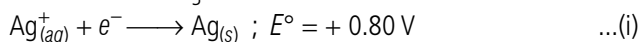
Oxidation :**Reduction :****Overall reaction :**

Since the EMF for the above reaction is positive, therefore, the above reaction is feasible.

2. (i) In aqueous solution, AgNO_3 ionises to give $\text{Ag}_{(aq)}^{+}$ and $\text{NO}_{3(aq)}^{-}$ ions.

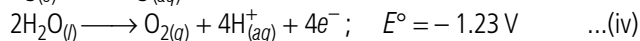
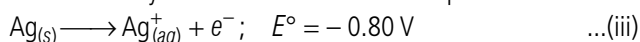


Thus, when electricity is passed $\text{Ag}_{(aq)}^{+}$ ions move towards the cathode while NO_3^{-} ions move towards the anode.



Since the electrode potential (*i.e.*, reduction potential of $\text{Ag}_{(aq)}^{+}$ ions is higher than that of H_2O molecules, therefore, at the cathode, $\text{Ag}_{(aq)}^{+}$ ions (rather than H_2O molecules) are reduced.

Similarly, at the anode, either Ag metal of the anode or H_2O molecules may be oxidised. Their electrode potentials are :



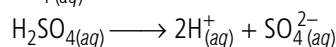
Since the oxidation potential of Ag is much higher than that of H_2O therefore, at the anode, Ag of the silver anode gets oxidised and not the H_2O molecules. It may, however, be mentioned here that the oxidation potential of NO_3^{-} ions is even lower than that of H_2O since more bonds are to be broken during reduction of NO_3^{-} ions than those in H_2O .

Thus, when an aqueous solution of AgNO_3 is electrolysed, Ag from Ag anode dissolves while $\text{Ag}_{(aq)}^{+}$ ions present in the solution gets reduced and gets deposited on the cathode.

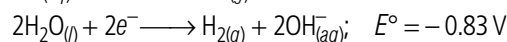
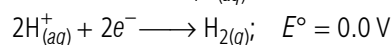
(ii) When electrolysis of AgNO_3 solution is carried out using platinum electrodes, instead of silver electrodes, oxidation of water occurs at the anode since Pt being a noble metal does

not undergo oxidation easily. As a result, O_2 is liberated at the anode according to equation (iv).

(iii) In aqueous solution, H_2SO_4 ionises to give $H_{(aq)}^+$ and SO_4^{2-} ions.



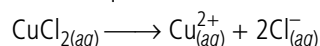
Thus, when electricity is passed, $H_{(aq)}^+$ ions move towards cathode while SO_4^{2-} ions move towards anode.



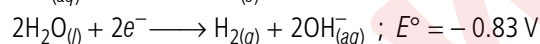
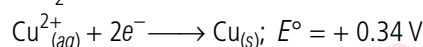
Since the electrode potential (*i.e.* reduction potential) of $H_{(aq)}^+$ ions is higher than that of H_2O , therefore, at the cathode, $H_{(aq)}^+$ ions (rather than H_2O molecules) are reduced to evolve H_2 gas. Similarly at the anode, either SO_4^{2-} ions or H_2O molecules are oxidised. Since the oxidation potential of SO_4^{2-} is expected to be much lower (since it involves cleavage of many bonds as compared to those in H_2O) than that of H_2O molecules, therefore, at the anode, it is H_2O molecules (rather than SO_4^{2-} ions) which are oxidised to evolve O_2 gas.

From the above discussion, it follows that during electrolysis of an aqueous solution of H_2SO_4 only the electrolysis of H_2O occurs liberating H_2 at the cathode and O_2 at the anode.

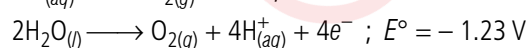
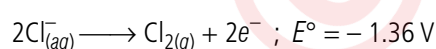
(iv) In aqueous solution, $CuCl_2$ ionises as follows :



On passing electricity, $Cu_{(aq)}^{2+}$ ions move towards cathode and $Cl_{(aq)}^-$ ions move towards anode. Thus, at cathode, either $Cu_{(aq)}^{2+}$ or H_2O molecules are reduced. Their electrode potentials are :



Since the electrode potential of $Cu_{(aq)}^{2+}$ ions is much higher than that of H_2O , therefore, at the cathode, $Cu_{(aq)}^{2+}$ ions are reduced and not H_2O molecules. Similarly, at the anode, either $Cl_{(aq)}^-$ ions or H_2O molecules are oxidized. Their oxidation potentials are :



Although oxidation potential of H_2O molecules is higher than that of Cl^- ions, nevertheless, oxidation of $Cl_{(aq)}^-$ ions occurs

in preference to H_2O since due to overvoltage of O_2 , much more potential than -1.36 V is required for the oxidation of chloride ions.

Thus, when an aqueous solution of $CuCl_2$ is electrolysed, Cu metal is liberated at the cathode while Cl_2 gas is evolved at the anode.

3. We know that,

$$E_{Al^{3+}/Al}^\circ = -1.66 \text{ V}, E_{Cu^{2+}/Cu}^\circ = +0.34 \text{ V},$$

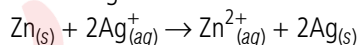
$$E_{Fe^{2+}/Fe}^\circ = -0.44 \text{ V}, E_{Mg^{2+}/Mg}^\circ = -2.36 \text{ V} \text{ and } E_{Zn^{2+}/Zn}^\circ$$

$$= -0.76 \text{ V}.$$

Since a metal with lower electrode potential is a stronger reducing agent, therefore, Mg can displace all the above metals from their aqueous solutions, Al can displace all metals except Mg from the aqueous solutions of their salts. Zn can displace all metals except Mg and Al from the aqueous solutions of their salts while Fe can displace only Cu from the aqueous solution of its salts. Thus, the order in which they can displace each other from the solution of their salts is Mg, Al, Zn, Fe, Cu .

4. Lower the electrode potential, better is the reducing power. Since the electrode potentials increase in the order; $K^+/K(-2.93 \text{ V})$, $Mg^{2+}/Mg(-2.37 \text{ V})$, $Cr^{3+}/Cr(-0.74 \text{ V})$, $Hg^{2+}/Hg(0.79 \text{ V})$, $Ag^+/Ag(0.80 \text{ V})$, therefore, reducing power of metals increases in the order, *i.e.*, $Ag < Hg < Cr < Mg < K$.

5. The given redox reaction is



Since Zn gets oxidized to Zn^{2+} ions, and Ag^+ gets reduced to Ag metal, therefore, oxidation occurs at zinc electrode and reduction occurs at the silver electrode.

Thus, galvanic cell corresponding to the above redox reaction may be depicted as : $Zn | Zn_{(aq)}^{2+} || Ag_{(aq)}^+ | Ag$

(i) Since oxidation occurs at the zinc electrode, therefore, electrons accumulate on the zinc electrode and hence, zinc electrode is negatively charged.

(ii) The ions carry current. The electrons flow from Zn to Ag electrode while the current flows from Ag to Zn electrode.

(iii) The reactions occurring at the two electrodes are :

