8 Redox Reactions

LESSON AT A GLANCE

- **Redox reaction:** Reactions in which oxidation and reduction occur simultaneously are called redox reactions.
- **Oxidation:** Involves loss of one or more electrons.
- **Reduction:** Involves gain of one or more electrons.
- **Oxidising agent:** Accepting electrons.
- **Reducing agent:** Losing electrons.
- **Electrochemical cell:** It is a device in which redox reaction is carried indirectly and decrease in energy gives electrical energy.
- **Electrode potential:** It is the potential difference between the electrode and its ions in solution.
- **Standard electrode potential:** It is the potential of an electrode with respect to standard hydrogen electrode.
- **Electrochemical series:** It is activity series. It has been formed by arranging the metals in order of increasing standard reduction potential value.

TEXTBOOK QUESTIONS SOLVED

Q1. *Assign oxidation number to the underlined elements in each of the following species:*

(a) NaH_2PO_4 (b) $NaHSO_4$

(c) $H_4P_2O_7$ (d) $K_2\underline{Mn}O_4$

(e) CaO_2 (f) $NaBH_4$

(g) $H_2S_2O_7$ (h) $KAl(SO_4)_2.12H_2O$

Ans. Following are the answers:

(a)

P in $\overset{+1}{Na}\,\overset{+1}{H_2}\,\overset{x}{P}\,\overset{-2}{O_4}$

$$(+1) + 2(+1) + x + 4(-2) = 0$$
$$x + 3 - 8 \text{ or } x = +5$$

(*b*) S in $NaHSO_4$

$\overset{+1}{Na}\ \overset{+1}{H}\ \overset{x}{S}\ \overset{-2}{O_4}$

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

(*c*) P in $H_4P_2O_7$

$\overset{+1}{H_4}\ \overset{x}{P_2}\ \overset{-2}{O_7}$

$$4(+1) + 2x + 7(-2) = 0$$
$$2x - 10 = 0$$
$$x = +5$$

(*d*) Mn in K_2MnO_4

$\overset{+1}{K_2}\ \overset{x}{Mn}\ \overset{-2}{O_4}$

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

(*e*) Let the oxidation number of CaO_2 be x.

$$2 \times 2x = 0 \quad (\because \text{oxy No. of } a = +2)$$
$$x = -1$$

Ca O_2

Thus, oxidation number of O in CaO_2 = **-1.**

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

$\overset{+1}{Na}\ \overset{x}{B}\ \overset{-1}{H_4} \quad \therefore \quad 1(+1) + x + 4(-1) = 0 \quad \text{or} \quad x = +3$

Thus, the oxidation number of B in $NaBH_4$ = **+ 3.**

(*g*) $\overset{+1}{Na_2}\ \overset{x}{S_2}\ \overset{-2}{O_7} \quad \therefore \quad 2(+1) + 2(x) + 7(-2) = 0 \quad \text{or} \quad x = +6$

Thus, the oxidation number of S in $Na_2S_2O_7$ = **+ 6.**

(*h*) $\overset{+1}{K}\ \overset{+3}{Al}\ (\overset{x}{S}\ \overset{-2}{O_4})_2\ 12(\overset{+1}{H_2}\ \overset{-2}{O}) \quad \text{or} \quad +1 + 3 + 2x + 8(-2)$

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

Alternatively, since H_2O is a neutral molecule, therefore, sum of oxidation numbers of all the atoms in H_2O may be taken as zero. As such water molecules may be ignored while computing the oxidation number of S.

$\therefore \quad +1+3+2x-16=0$ or $x=+6$

Thus, the oxidation number of S in $KAl(SO_4)_2.12H_2O$ = **+6.**

Q2. *What are the oxidation numbers of the underlined elements in each of the following and how do you rationalise your results?*

(a) KI_3 (b) $H_2S_4O_6$

(c) Fe_3O_4 (d) CH_3CH_2OH

(e) CH_3COOH

Ans. (*a*) Using rules to determine oxidation number, the oxidation number of I in KI_3 is –1/3. Since the fractional oxidation number cannot be the possibility and investigation of the struture of KI_3 or I_3^- should be done. KI_3 exists as $K^+[I_2 \leftarrow I^-]$ where a dative bond exists between I_2 molecule and I^- ion. Oxidation number of molecular iodine, I_2, is to be taken as zero. The oxidation numbers are:

$$\overset{+1}{K}[\overset{0}{I_2} \leftarrow \overset{-1}{I}]$$

Thus, the oxidation numbers as obtained would be, K = +1, I in I_2 = 0, I = –1. Since there are three iodine atoms, the average oxidation number per atom would be –1/3.

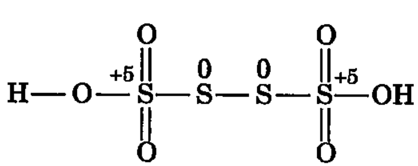
(*b*) *By conventional method.* O.N. of S in $H_2S_4O_6$ =

$$\overset{+1}{H_2}\ \overset{x}{S_4}\ \overset{-2}{O_6}$$

or $2(+1)+4x+6(-2)=0$ or $x=+2.5$ (*wrong*)

But it is wrong because all the four S atoms cannot be in the same oxidation state.

By chemical bonding method. The structure of $H_2S_4O_6$ is shown below:



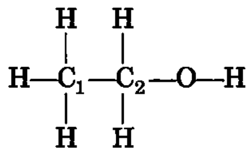
The O.N. of each of the S-atoms linked with each other in the middle is zero while that of each of the remaining two S-atoms is +5.

(*c*) *By conventional method.* O.N. of Fe in $\overset{x}{Fe_3}\overset{-2}{O_4}$ or $3x + 4(-2) = 0$ or $x = 8/3$.

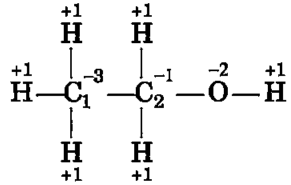
By stoichiometry. $Fe_3O_4 \equiv \overset{+2}{Fe}\overset{-2}{O} \cdot \overset{+3}{Fe_2}\overset{-2}{O_3}$.

$\therefore$ O.N. of Fe in Fe_3O_4 + 2 and + 3

(*d*) CH_3CH_2OH. For the oxidation number of the two C atoms, the structure of the molecule should be written as:



If oxidation number of C atom is calculated by using the formula C_2H_6O, it is –2. From the structure, carbon atom labelled as C_2 will have the oxidation number as –1 and on C_1 it is –3.

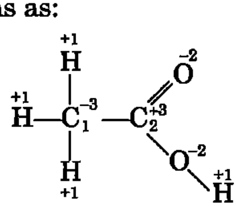


Since the average oxidation number is –2, which may be obtained as $\frac{-3+(-1)}{2} = -2$.

So, the oxidation number on C_1 is –3 and C_2 is –1.

(*e*) Oxidation number of C atom in CH_3COOH or $C_2H_4O_2$ is zero.

From the structure, we obtain oxidation numbers of two carbon atoms as:



Oxidation number of C_2 is +3 while that of C_1 is –3. The average oxidation number is (+3 – 3) ÷ 2 = 0.

Q3. *Justify that the following reactions are redox reactions:*

(*a*) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$

(*b*) $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$

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

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

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

Ans. (*a*) $\overset{+2-2}{CuO}(s) + \overset{0}{H_2}(g) \longrightarrow \overset{0}{Cu}(s) + \overset{+1-2}{H_2O}(g)$

Here, *O* is removed from CuO, therefore, it is reduced to Cu while *O* is added to H_2 to form H_2O, therefore, it is oxidised. Further, O.N. of Cu decreases from + 2 in CuO to 0 in Cu but that of *H* increases from 0 in H_2 to +1 in H_2O. Therefore, CuO is reduced to Cu but H_2 is oxidised to H_2O. Thus, this is a redox reaction.

(*b*) $\overset{+3}{Fe_2}\overset{-2}{O_3}(s) + 3\overset{+2}{C}O(g) \longrightarrow 2\overset{0}{Fe}(s) + 3\overset{+4}{C}O_2(g)$

Here O.N. of Fe decreases from +3 if Fe_2O_3 to 0 in Fe while that of *C* increases from +2 in CO to +4 in CO_2. Further, oxygen is removed from Fe_2O_3 and added to CO, therefore, Fe_2O_3 is reduced while CO is oxidised. Thus, this is a redox reaction.

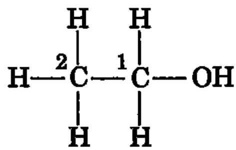
(*c*) $4\overset{+3-1}{BCl_3}(g) + \overset{+1\,+3\,-1}{LiAlH_4}(s) \longrightarrow 2\overset{-3}{B_2}\overset{+1}{H_6}(g) + 3\overset{+1\,-1}{LiCl}(s) + 3\overset{+3}{Al}\overset{-1}{Cl_3}(s)$

Here, O.N. of *B* decreases from +3 in $BrCl_3$ to -3 in B_2H_6 while that of *H* increases from –1 in $LiAlH_4$ to +1 in B_2H_6. Therefore, BCl_3 is reduced while $LiAlH_4$ is oxidised. Further, *H* is added to BCl_3 but is removed from $LiAlH_4$, therefore, BCl_3 is reduced while $LiAlH_4$ is oxidised. Thus, it is a redox reaction.

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

Here, each *K* atom as lost one electron to form K^+ while F_2 has gained two electrons to form two F^- ions. Therefore, *K* is oxidised while F_2 is reduced. Thus, it is a redox reaction.

By chemical bonding, C_2 is attached to three *H*-atoms (less electronegative than carbon) and one CH_2OH group (more electronegative than carbon), therefore,



O.N. of $C_2 = 3(+1) + x + 1(-1) = 0$ or $x = -2$

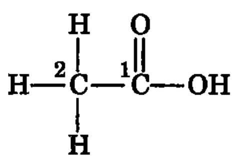
C_1 is, however, attached to one OH (O.N. = –1) and one CH_3 (O.N. = +1) group.

Therefore, O.N. of $C_1 = +1 + 2(+1) + x + 1(-1) = 0$ or $x = -2$

(*e*) *By conventional method.* $CH_3COOH = \overset{x}{C_2}\overset{+1}{H_4}\overset{-2}{O_2}$ or

$2x + 4 - 4 = 0$ or $x = 0$

By chemical bonding method, C_2 is attached to three *H*-atoms (less electronegative than carbon) and one—COOH group (more electronegative than carbon,



therefore, O.N. of $C_2 = 3(+1) + x + 1(-1) = 0$ or $x = -2$

C_1 is, however, attached to one oxygen atom by a double bond, one OH (O.N. = –1) and one CH_3 (O.N. = +1) group, therefore, O.N. of $C_1 = +1 + x + 1(-2) + 1(-1) = 0$ or $x = +2$

Q4. *Fluorine reacts with ice and results in the change:*

$$H_2O\,(s) + F_2\,(g) \longrightarrow HF\,(g) + HOF\,(g)$$

Justify that this reaction is a redox reaction.

Ans.

$$\overset{+1\ -2}{H_2O}\,(s) + \overset{0}{F_2}\,(g) \longrightarrow \overset{+1\ -1}{H\ F}\,(g) + \overset{+1\ -2\ +1}{H\ O\ F}\,(g)$$

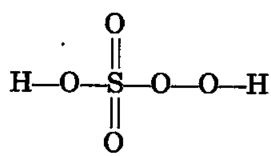
Q5. *Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO, $Cr_2O_7^{2-}$ and NO_3^-. Suggest structures of these compounds. Count for the fallacy.*

Ans. (*a*) Oxidation number of S atom in H_2SO_5 is

$$\overset{+1}{H_2}\overset{x}{S}\overset{-2}{O_5};\ 2(+1) + x + 5(-2) = 0$$

Therefore, $x = +10 - 2 = +8$

As the valence shell of S cannot have more than six electrons thus the oxidation number of S cannot be more than six. This fallacy can be removed by seeing the structure of H_2SO_5 (peroxomonosulphuric acid):



Here, sulphur atom is attached to two oxygen atoms, one OH group and one peroxo group (O—O—H). The oxidation number of OH is –1 and that of (O—O—H) group is –1 $(\overset{-2}{O_2}\overset{+1}{H})$. We can obtain oxidation number of S as:

–1	+	x	+	2(–2)	+	(–1) = 0
(from OH)				from two O atoms		from one O_2H group

$\therefore \qquad x = +1 + 1 + 4 = +6$

So, the oxidation number of sulphur is +6. This can also be achieved by writing the formula from the structure as $H_2S(O_2)O_3$ and obtaining the oxidation number as:

$$\overset{+1}{H_2}\overset{x}{S}\,(\overset{-1}{O_2})\,\overset{-2}{O_3}$$

$$\therefore \quad 2(+1) + x + 2(-1) + 3(-2) = 0$$

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

$$x = 6 + 2 - 2$$

$$= 6 + 0 = 6$$

(*b*) Oxidation number of Cr in $Cr_2O_7^{2-}$ is

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

So, $\quad x = (-2 + 14) \div 2 = +6$

There is no fallacy.

(*c*) Oxidation number of N in NO_3^-

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

So, $\quad x = -1 + 6 = +5$

In this case, there is no fallacy since the oxidation number of N cannot exceed five.

Q6. *Write formulas for the following compounds:*

(*a*) *Mercury (II) chloride*
(*b*) *Nickel (II) sulphate*
(*c*) *Tin (IV) oxide*
(*d*) *Thallium (I) sulphate*
(*e*) *Iron (III) sulphate*
(*f*) *Chromium (III) oxide*

Ans. (*a*) $HgCl_2$ or $Hg^{II}Cl_2$
(*b*) $NiSO_4$ or $Ni^{II}SO_4$
(*c*) SnO_2 or $Sn^{IV}O_2$
(*d*) Tl_2SO_4 or $(Tl^{I})_2SO_4$
(*e*) $Fe_2(SO_4)_3$ or $(Fe^{III})_2(SO_4)_3$
(*f*) Cr_2O_3 or $(Cr^{III})_2O_3$

Q7. *Suggest a list of the substances where carbon can exhibit oxidation states from –4 to +4 and nitrogen from –3 to +5.*

Ans.

Compound of Carbon	Ox. No. of C	Compound of Nitrogen	Ox. No. of N
CH_4	–4	NH_3 or NCl_3	–3
C_2H_6	–3	N_2H_4	–2
CH_3Cl	–2	KNH_2	–1
C_2H_2	–1	N_2	0
		N_2O	+1
CH_2Cl_2	0	NO	+2
$CHCl_3$	+2	N_2O_3	+3
(also CO)		(also HNO_2)	
		N_2O_4	+4
CCl_4	+4	N_2O_5	+5
(also CO_2)		(also HNO_3)	

Q8. *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?*

Ans. (*i*) Oxidation number of sulphur can range from –2 to +6. In SO_2, the oxidation number is +4, so in reactions its oxidation number can decrease or increase in the products. Hence, SO_2 can act both as an oxidising agent or reducing agent.

(*ii*) In H_2O_2, the oxidation number of oxygen is –1. However, oxidation number of oxygen can vary from –2 to 0 or +2 only in OF_2. Hence, H_2O_2 can act both as oxidising or reducing agent.

(*iii*) In the case of ozone, O_3, the oxidation of oxygen is zero which can only decrease to either 0, –1, or –2. So, it acts only as an oxidising agent.

(*iv*) In HNO_3, nitrogen has its maximum oxidation number (+5) which can only decrease and not increase. So, HNO_3 acts only as an oxidising agent.

Q9. *Consider the reactions:*

(*a*) $6CO_2 (g) + 6H_2O (l) \longrightarrow C_6H_{12}O_6 (aq) + 6O_2 (g)$

(*b*) $O_3 (g) + H_2O_2 (l) \longrightarrow H_2O (l) + 2O_2 (g)$

Why it is more appropriate to write these reactions as:

(*a*) $6CO_2 (g) + 12H_2O (l) \longrightarrow C_6H_{12}O_6 (aq) + 6H_2O (l) + 6O_2 (g)$

(*b*) $O_3 (g) + H_2O_2 (l) \longrightarrow H_2O (l) + O_2 (g) + O_2(g)$

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

Ans. (*a*) To understand that why is the photosynthesis reaction written as:

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

we should understand the mechanism of photosynthesis reaction.

Photosynthesis is a remarkable redox reaction where conversion of water and carbon dioxide takes place into carbohydrate and oxygen using solar energy. Isotope tracer technique using ^{18}O isotope has shown that production of carbohydrate involve the *reduction* of CO_2 and the oxidation of H_2O to O_2. This redox reaction takes place in the presence of chlorophyll, the green matter of plants.

The two processes in a simplified way can be written as:

Oxidation of H_2O :

$$12H_2O (l) \longrightarrow 6O_2 (g) + 24 [H]$$

Reduction of CO_2:

$6CO_2\,(g) + 24[H] \longrightarrow C_6H_{12}O_6\,(aq) + 6H_2O\,(l)$

Here, 24[H] does not imply free atoms of hydrogen but represents their reducing capacity.

Adding the two reactions and cancelling out 24[H] to get the required equation:

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

This reaction shows that 12 moles of water molecules are used to form 1 mole of $C_6H_{12}O_6$ and in this process, 6 moles of H_2O molecules are also formed. However, when H_2O molecules on both the sides are balanced, a simple reaction results:

$$6CO_2\,(g) + 6H_2O\,(l) \xrightarrow[\text{chlorophyll}]{h\nu} C_6H_{12}O_6\,(aq) + 6O_2\,(g)$$

(*b*) For the simple reaction:

$O_3\,(g) + H_2O_2\,(l) \longrightarrow H_2O\,(l) + 2O_2\,(g)$

The more appropriate reaction that is written as:

$O_3\,(g) + H_2O_2\,(l) \longrightarrow H_2O\,(l) + O_2\,(g) + O_2\,(g)$

can be understood by splitting the redox reaction as:

$$O_3\,(g) \longrightarrow O_2\,(g) + \frac{1}{2}O_2\,(g)$$

$$H_2O_2\,(l) + \frac{1}{2}O_2\,(g) \longrightarrow H_2O\,(l) + O_2\,(g)$$

$$\overline{O_3\,(g) + H_2O_2\,(l) \longrightarrow H_2O\,(l) + O_2\,(g) + O_2\,(g)}$$

The mechanism or path of this reaction can also be determined by using isotope tracer technique.

Q10. *The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why?*

Ans. AgF_2 besides being a strong oxidising agent is also a fluorinating agent. It decomposes on heating giving out F_2 gas because of which it is a strong oxidising agent.

$$2AgF_2 \xrightarrow{\text{heat}} 2AgF + F_2$$

Also, the oxidation number of Ag changes from +2 to +1 and in this regard also it acts as an oxidising agent. Thus, the combined effect of change in oxidation number and release of F_2 gas on decomposition makes AgF_2 a strong oxidising agent.

Q11. *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.*

Ans. The answer to this problem can be had when an oxidising agent forms different compounds with the reducing agent.

(*a*) Reaction between carbon (reducing agent) and oxygen (an oxidising agent)

$$\underset{\text{(Excess)}}{2\overset{0}{C}\,(s)} + O_2\,(g) \longrightarrow 2\overset{+2}{C}O\,(g)$$

$$\overset{0}{C}\,(s) + \underset{\text{(Excess)}}{O_2\,(g)} \longrightarrow \overset{+4}{C}O_2\,(g)$$

(*b*) Reaction between phosphorus and oxygen:

$$\underset{\text{(Excess)}}{P_4\,(s)} + 3O_2\,(g) \longrightarrow \overset{+3}{P_4}O_6\,(s)$$

$$P_4\,(s) + \underset{\text{(Excess)}}{5O_2\,(g)} \longrightarrow \overset{+5}{P_4}O_{10}\,(s)$$

(*c*) Reaction between sulphur and oxygen:

$$\underset{\text{(Excess)}}{S_8\,(s)} + 8O_2\,(g) \longrightarrow 8\overset{+4}{S}O_2\,(g)$$

$$S_8\,(s) + \underset{\text{(Excess)}}{12O_2\,(g)} \xrightarrow{\text{catalyst}} 8\overset{+6}{S}O_3\,(g)$$

Q12. *How do you count for the following observations?*

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

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

Ans. (*a*) [Note: In the question, it is stated that *alcoholic* $KMnO_4$ is used as an oxidant to oxidise toluene, $C_6H_5CH_3$, to benzoic acid, C_6H_5COOH. Probably, it is *alkaline* $KMnO_4$ and not alcoholic otherwise alcohol, C_2H_5OH, will also be oxidised to acetic acid, CH_3COOH. With alkaline $KMnO_4$, the redox reaction is

$$[2KMnO_4\,(aq) + 2KOH\,(aq) \longrightarrow 2K_2MnO_4\,(aq) + H_2O\,(l) + (O)] \times 3$$

$$C_6H_5CH_3\,(l) + 3(O) \longrightarrow C_6H_5COOH\,(s) + H_2O\,(l)$$

$$C_6H_5COOH\,(s) + KOH\,(aq) \longrightarrow C_6H_5COOK\,(aq) + H_2O\,(l)$$

$$6KMnO_4\,(aq) + 7KOH\,(aq) + C_6H_5CH_3\,(l) \longrightarrow 6K_2MnO_4\,(aq) + C_6H_5COOK\,(aq) + 5H_2O\,(l)$$

C_6H_5COOK is water soluble and to regenerate C_6H_5COOH an acid, such as H_2SO_4 is added.

(*b*) HCl is produced by the reaction:

$$2NaCl\,(s) + 2H_2SO_4\,(l) \longrightarrow 2NaHSO_4\,(s) + 2HCl\,(g)$$

Here, a stronger acid H_2SO_4 displaces a weaker one. However, further reaction with HCl (or Cl^- ions) does not take place because Cl^- ions or HCl cannot reduce H_2SO_4 to SO_2 (g) as:

$$H_2SO_4\,(l) + 2HCl\,(g) \longrightarrow Cl_2\,(g) + SO_2\,(g) + 2H_2O\,(l)$$

On the other hand, HBr, as produced by the action of H_2SO_4 on NaBr, is a stronger reducing agent than HCl and so reduces H_2SO_4 to SO_2 (g) and red vapours of Br_2 (g):

$$2H_2SO_4\,(l) + 2NaBr\,(s) \longrightarrow 2NaHSO_4\,(s) + 2HBr\,(g)$$

$$2HBr\,(g) + H_2SO_4\,(l) \longrightarrow Br_2\,(g) + SO_2\,(g) + 2H_2O\,(l)$$

Q13. *Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:*

(*a*) $2AgBr\,(s) + C_6H_6O_2\,(aq) \longrightarrow 2Ag\,(s) + 2HBr\,(aq) + C_6H_4O_2\,(aq)$

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

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

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

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

Ans.

	Substance oxidised	Substance reduced	Oxidising agent	Reducing agent
(a)	$C_6H_6O_2 (aq)$	AgBr (s)	AgBr (s)	$C_6H_4O_2 (aq)$
(b)	HCHO (l)	$[Ag(NH_3)_2]^+ (aq)$	$[Ag(NH_3)_2]^+ (aq)$	HCHO (l)
(c)	HCHO (l)	$Cu^{2+} (aq)$	$Cu^{2+} (aq)$	HCHO (l)
(d)	$N_2H_4 (l)$	$H_2O_2 (l)$	$H_2O_2 (l)$	$N_2H_4 (l)$
(e)	Pb (s)	$PbO_2 (s)$	$PbO_2 (s)$	Pb (s)

Q14. *Consider the reactions:*

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

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

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

Ans. In both the equations, $S_2O_3^{2-}$ ions are reducing agents while I_2 and Br_2 are oxidising agents. Br_2 is a stronger oxidising agent than I_2. A stronger oxidising agent is capable of bringing about the higher oxidation state or oxidation number of the reducing agent, *i.e.*,

(i) $\overset{+2}{S_2}O_3^{2-} \xrightarrow{I_2} \overset{2.5}{S_4}\overset{2-}{O_6^{2-}}$

(2.5 is the average o. n. of S atom)

(ii) $\overset{+2}{S_2}O_3^{2-} \xrightarrow{Br_2} \overset{+6}{S}O_4^{2-}$

Thus, Br_2 and I_2 react with $S_2O_3^{2-}$ differently.

Q15. *Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.*

Ans. (a) From the E^0_{red} values of X^-/X_2 couples involving halogens, it can be seen that the oxidising power decreases down the group:

$F_2 > Cl_2 > Br_2 > I_2$

F_2, Cl_2 and Br_2 are capable of displacing the less oxidising halogens from their halides according to the equation:

$$2MX + Y_2 \longrightarrow 2MY + X_2$$

(Y_2 is a stronger oxidising agent than X_2)

I_2 will not be able to displace any halogen because it is least oxidising. With F_2 a reaction with H_2O also takes place:

$$H_2O\ (l) + F_2\ (g) \longrightarrow HF\ (aq) + O_2\ (g)$$

(*b*) Among hydrohalic compounds, the reducing power decreases as:

HI > HBr > HCl > HF

It can also be seen from the E^0_{red} values. HBr and HI both can reduce H_2SO_4 to SO_2 and their respective halogens as:

$$H_2SO_4\ (l) + 2HI\ (aq) \longrightarrow I_2\ (g) + SO_2\ (g) + 2H_2O\ (l)$$

Q16. *Why does the following reaction occur?*

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

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

Ans. The given reaction:

$$\overset{+8}{XeO_6^{4-}}\ (aq) + \overset{-1}{2F^-}\ (aq) + 6H^+\ (aq) \longrightarrow \overset{+6}{XeO_3}\ (g) + \underset{0}{F_2}\ (g) + 3H_2O\ (l)$$

occurs because XeO_6^{4-} ions are stronger oxidising agent than F_2 (g) as can be seen from the change in oxidation numbers. F^- ions are oxidised to F_2 while XeO_6^{4-} ions are reduced to XeO_3.

The compound Na_4XeO_6 is a stronger oxidising agent than F_2 gas.

Q17. *Consider the reactions:*

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

(*b*) $H_3PO_2\ (aq) + 2CuSO_4\ (aq) + 2H_2O\ (l) \longrightarrow H_3PO_4\ (aq) + 2Cu\ (s) + H_2SO_4\ (aq)$

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

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

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

Ans. (*a*) Ag^+ ions are acting as oxidising agent and $H_2PO_2^-$ ions as reducing agent, *i.e.*, $AgNO_3$ is an oxidising agent and H_3PO_2 is a reducing agent.

(*b*) H_3PO_2 is a reducing agent while $CuSO_4$ (or Cu^{2+} ions) is oxidising agent.

(*c*) $[Ag(NH_3)_2]^+$ is an oxidising agent and C_6H_5CHO is a reducing agent.

(*d*) Redox reaction is not taking place. Thus, Cu^{2+} ions or C_6H_5CHO are neither acting as reducing agent nor oxidising agent.

Q18. *Balance the following redox reactions by ion – electron method:*

(*a*) $MnO_4^- (aq) + I^- (aq) \longrightarrow MnO_2 (s) + I_2 (s)$ (*in basic medium*)

(*b*) $MnO_4^- (aq) + SO_2 (g) \longrightarrow Mn^{2+} (aq) + HSO_4^- (aq)$ (*in acidic solution*)

(*c*) $H_2O_2 (aq) + Fe^{2+} (aq) \longrightarrow Fe^{3+} (aq) + H_2O (l)$ (*in acidic solution*)

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

Ans. (*a*) Do it yourself.

(*b*) The balanced half reaction equations are:

Oxidation half equation:

$$SO_2(g) + 2H_2O(l) \longrightarrow HSO_4^-(aq) + 3H^+(aq) + 2e^- \quad ...(i)$$

Reduction half equation:

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

Multiply Eq. (*i*) by 5 and Eq. (*ii*) by 2 and add, we have,

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

(*c*) *Oxidation half equation:*

$$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^- \quad ...(i)$$

Reduction half equation:

$$H_2O_2(aq) + 2H^+(aq) + 2e^- \longrightarrow 2H_2O(l) \quad ...(ii)$$

Multiply Eq. (*i*) by 2 and add it to Eq. (*ii*), we have,

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

(*d*) Following the procedure detailed, the balanced half reaction equations are:

Oxidation half equation:

$SO_2(g) + 2H_2O(l) \longrightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$ ...(*i*)

Reduction half equation:

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

Multiply Eq. (*i*) by 3 and add it to Eq. (*ii*), we have,

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

Q19. *Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.*

(*a*) $P_4\,(s) + OH^-\,(aq) \longrightarrow PH_3\,(g) + HPO_2^-\,(aq)$

(*b*) $N_2H_4\,(l) + ClO_3^-\,(aq) \longrightarrow NO\,(g) + Cl^-\,(g)$

(*c*) $Cl_2O_7\,(g) + H_2O_2\,(aq) \longrightarrow ClO_2^-\,(aq) + O_2\,(g) + H^+$

Ans. (*a*) $\overset{0}{P_4}(s) + OH^-(aq) \longrightarrow \overset{-3}{P}\overset{+1}{H_3}(g) + \overset{+1}{H_2}\overset{+1}{P}\overset{-2}{O_2^-}$

O.N. increases by 1 per P atom.

P_4 acts both as an oxidising as well as a reducing agent.

Oxidation number method:

Total decrease in O.N. of P_4 in $PH_3 = 3 \times 4 = 12$

Total increase in O.N. of P_4 in $H_2PO_2^- = 1 \times 4 = 4$

Therefore, to balance increases decreases in O.N. multiply PH_3 by 1 and $H_2PO_2^-$ by 3, we have,

$$P_4(s) + OH^-(aq) \longrightarrow PH_3(g) + 3H_2PO_2^-(aq)$$

To balance O atoms, multiply OH^- by 6, we have,

$$P_4(s) + 6OH^-(aq) \longrightarrow PH_3(g) + 3H_2PO_2^-(aq)$$

To balance H atoms, add $3H_2O$ to L.H.S. and $3OH^-$ to the R.H.S., we have,

$$P_4(s) + 6OH^-(aq) + 3H_2O(l) \longrightarrow PH_3(g) + 3H_2PO_2^-(aq) + 3OH^-(aq)$$

or $P_4(s) + 3OH^-(aq) + 3H_2O(l) \longrightarrow PH_3(g) + 3H_2PO_2^-(aq)$ ...(*i*)

Thus, Eq. (*i*) represents the correct balanced equation.

Ion electron method. The two half reactions are:

Oxidation half reaction:

$$P_4(s) \longrightarrow H_2PO_2^-(aq) \qquad ...(ii)$$

Balancing *P* atoms, we have,

$$\overset{0}{P_4}(s) \longrightarrow 4H_2\overset{+1}{P}O_2^-(aq)$$

Balance O.N. by adding electrons,

$$P_4(s) \longrightarrow 4H_2PO_2^-(aq) + 4e^-$$

Balance charge by adding 8 OH^- ions,

$$P_4(s) + 8OH^-(aq) \longrightarrow 4H_2PO_2^-(aq) + 4e^- \qquad ...(iii)$$

O and H get automatically balanced. Thus, Eq. (*iii*) represents the balanced oxidation half reaction.

Reduction half reaction:

$$\overset{0}{P_4}(s) \longrightarrow \overset{-3}{P}H_3(g) \qquad ...(iv)$$

Balancing P atoms, we have,

$$P_4(s) \longrightarrow 4PH_3(g)$$

Balance O.N. by adding electrons,

$$P_4(s) + 12e^- \longrightarrow 4PH_3(g)$$

Balance charge by adding 12OH^- ions,

$$P_4(s) + 12e^- \longrightarrow 4PH_3(g) + 12OH^-(aq)$$

Balance *O* atoms, by adding 12H_2O to L.H.S. of above equation.

$$P_4(s) + 12H_2O(l) + 12e^- \longrightarrow 4PH_3(g) + 12OH^-(aq) \qquad ...(v)$$

To cancel out electrons, multiply Eq. (*iii*) by 3 and add it to Eq. (*v*), we have,

$$4P_4(s) + 24OH^-(aq) + 12H_2O(l) \longrightarrow$$
$$4PH_3(aq) + 12H_2PO_2^-(aq) + 12H_2O(l) + 12OH^-(aq)$$

or $P_4(g) + 3OH^-(aq) + 3H_2O(l) \longrightarrow$

$$PH_3(aq) + 3H_2PO_2^-(aq) \qquad ...(i)$$

Thus, Eq. (*i*) represents the correct balanced equation.

O.N. increases by 1 per P atom.

$$(b)\ \overset{0}{P_4}(s) + OH^-(aq) \longrightarrow \overset{-3}{P}\overset{+1}{H_3}(g) + \overset{+1}{H_2}\overset{+1}{P}\overset{-2}{O_2^-}$$

O.N. increases by 1 per P atom.

Therefore, N_2H_4 acts as the reducing agent while ClO_3^- acts as the oxidising agent.

Oxidation number method

Total increase in O.N. of N = 2 × 4 = 8

Total decreases in O.N. of Cl = 1 × 6 = 6

Therefore, to balance increase/decrease in O.N. multiply N_2H_4 by 3 and ClO_3^- by 4, we have,

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

To balance *N* and Cl atoms, multiply NO by 6 and Cl^- by 4, we have,

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

Balance O atoms by adding $6H_2O$,

$$3N_2H_4(l) + 4ClO_3^-(aq) \longrightarrow 6NO(g) + 4Cl^-(aq) + 6H_2O(l) \quad ...(i)$$

H atoms get automatically balanced and thus Eq. (*i*) represents the correct balanced equation.

Ion electron method.

Oxidation half equation:

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

Balance *N* atoms, $N_2H_4(l) \longrightarrow 2NO(g)$

Balance O.N. by adding electrons,

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

Balance charge by adding $8OH^-$ ions,

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

Balance O atoms by adding 6 H_2O,

$$N_2H_4(l) + 8OH^-(aq) \longrightarrow 2NO(g) + 6H_2O(l) + 8e^- \quad ...(ii)$$

Thus, Eq. (*ii*) represents the correct balanced oxidation half equation.

Reduction half equation

$$\overset{+5}{Cl}O_3^-(aq) \longrightarrow \overset{-1}{Cl^-}(aq)$$

Balance O.N. by adding electrons,

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

Balance charge by adding OH^- ions,

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

Balance O atoms by adding $3H_2O$,

$$ClO_3^-(aq) + 3H_2O(l) + 6e^- \longrightarrow Cl^-(aq) + 6OH^-(aq) \quad ...(iii)$$

Thus, Eq. (*iii*) represents the correct balanced reduction half equation.

To cancel out electrons gained and lost, multiply Eq. (*ii*) by 3 and Eq. (*iii*) by 4 and add, we have,

$$3N_2H_4(l) + 4ClO_3^-(aq) \longrightarrow 6NO(g) + 4Cl^-(aq) + 6H_2O(l) \quad ...(i)$$

Thus, Eq. (*i*) represents the correct balanced equation

O.N. decreases by 4 per Cl atom

$$(c)\ \overset{+7}{Cl_2}O_7(g) + H_2\overset{-1}{O_2}(aq) \longrightarrow \overset{+3}{ClO_2^-}(g) + \overset{0}{O_2}(g) + H^+$$

O.N. increases by 1 per O atom

Thus, $Cl_2O_7(g)$ acts as an oxidising agent while $H_2O_2(aq)$ as the reducing agent.

Oxidation number method

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

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

∴ To balance increase/decrease in O.N. multiply H_2O_2 and O_2 by 4, we have,

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

To balance Cl atoms, multiply ClO_2^- by 2, we have,

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

To balance O atoms, add $3H_2O$ R.H.S., we have,

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

To balance H atoms, add $2H_2O$ to R.H.S. and $2OH^-$ to L.H.S., we have,

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

This represents the balanced redox equation.

Ion electron method

Oxidation half equation:

$$H_2\overset{-1}{O_2}(aq) \longrightarrow \overset{0}{O_2}(g)$$

Balance O.N. by adding electrons,

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

Balance charge by adding $2OH^-$ ions,

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

Balance O atoms by adding $2H_2O$,

$$H_2O_2(aq) + 2OH^-(aq) \longrightarrow O_2(g) + 2H_2O(l) + 2e^- \quad ...(i)$$

Reduction half equation:

$$\overset{+7}{Cl_2}O_7(g) \longrightarrow \overset{+3}{Cl}O_2^-(aq)$$

Balance Cl atoms; $Cl_2O_7(g) \longrightarrow 2ClO_2^-(aq)$

Balance O.N. by adding electrons,

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

Add $6OH^-$ ions to balance charge:

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

Balance O atoms by adding $3H_2O$ to L.H.S., we have,

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

To cancel out electrons, multiply Eq. (*i*) by 4 and add it to Eq. (*ii*), we have,

$$4H_2O_2(aq) + 8OH^-(aq) + Cl_2O_7(g) + 3H_2O(l) \longrightarrow 2ClO_2^- + (aq) + 6OH^-(aq) + 4O_2(g) + 8H_2O(l)$$

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

Q20. *What sorts of information can you draw from the following reaction?*

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

Ans. The given reaction is

$$(\overset{+3}{C}\overset{-3}{N})_2(g) + 2OH^-(aq) \rightarrow \overset{+2}{C}\overset{-3}{N}^-(aq) + (\overset{+4}{C}\overset{-3}{N}\overset{-2}{O})^-(aq) + H_2O(l)$$

(*i*) It is a redox reaction.

(*ii*) It is a disproportionation reaction because the oxidation number of C decreases from +3 in $(CN)_2$ to +2 in CN^- and increases to +4 in CNO^-.

(*iii*) $(CN)_2$, cyanogen, gets simultaneously oxidised to CNO^- ion and reduced to CN^- ion.

Q21. *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.*

Ans. According to the given conditions, the unbalanced equation is:

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

Since the given reaction is a disproportionation reaction, Mn^{3+} ions will act both as an oxidising agent and also a reducing agent. With these parameters following half reactions will occur:

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

(Mn^{3+} as an oxidising agent)

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

(Mn^{3+} as a reducing agent)

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

which is a balanced equation.

Q22. *Consider the elements: Cs, Ne, I and F*

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

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

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

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

Ans. (*a*) F, fluorine, is the only element that exhibits only negative state because it is the most electronegative element.

(*b*) Cs, Caesium, is the only element showing the positive oxidation state because it is least electronegative among Group 1 elements. Also because of its larger size, its ionisation enthalpy is quite low which facilitates the formation of Cs^+ ions.

(*c*) I, iodine, shows both –ve and +ve oxidation states. It easily forms I^-, iodide ions and +ve oxidation states in compounds IF, IF_3, IF_5 and IF_7, respectively.

(*d*) Ne, neon, shows neither +ve nor –ve oxidation states because it is an inert gas with highly stable electronic configuration.

Q23. *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.*

Ans. The required unbalanced equation is:

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

To get a balanced equation, it is split into half-reactions, oxidation and reduction, and adding the two equations while cancelling out the electrons.

$$Cl_2\,(aq) + 2e^- \longrightarrow 2Cl^-\,(aq);\ \text{(reduction)}$$

$$SO_2\,(aq) + 2H_2O\,(l) \longrightarrow SO_4^{2-}\,(aq) + 4H^+\,(aq) + 2e^-;\ \text{(oxidation)}$$

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

This is the balanced equation.

Q24. *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.*

Ans. (*a*) There can be a number of non-metals that show disproportionation reactions, *e.g.*, P_4, Cl_2 and S_8.

(*i*) $P_4\,(s) + 3OH^-\,(aq) + 3H_2O\,(l) \longrightarrow PH_3\,(g) + 3H_2PO_2^-\,(aq)$

(*ii*) $Cl_2\,(aq) + 2OH^-\,(aq) \longrightarrow Cl^-\,(aq) + OCl^-\,(aq) + H_2O\,(l)$

(*iii*) $S_8\,(s) + 12OH^-\,(aq) \longrightarrow 4S^{2-}\,(aq) + 2S_2O_3^{2-}\,(aq) + 6H_2O\,(l)$

(*b*) The metals are: Cu, Ga and In in their +1 oxidation states.

(*i*) $2Cu^+\,(aq) \longrightarrow Cu^{2+}\,(aq) + Cu\,(s)$

(*ii*) $3Ga^+\,(aq) \longrightarrow Ga^{3+}\,(aq) + 2Ga\,(s)$

(*iii*) $3In^+\,(aq) \longrightarrow In^{3+}\,(aq) + 2In(s)$

Q25. *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?*

Ans. The required balanced equation is:

$$4NH_3\,(g) + 5O_2\,(g) \xrightarrow{\text{catalyst}} 4NO\,(g) + 6H_2O\,(g)$$

4×17 g 5×32 g 4×30 g

$= 68$ g $= 160$ g $= 120$ g

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

$\therefore$ 20 g of O_2 will react with $\frac{68}{160} \times 20$ g of NH_3

$= 8.5$ g of NH_3

In the question, 10 g of NH_3 is taken which is more than 8.5 g of NH_3. This means that NH_3 (g) is present in excess to the stoichiometric amount, therefore, the calculations are to be done with the amount of O_2 which is the limiting reagent.

Thus, from the equation,

160 g O_2 produces 120 g NO

$\therefore$ 20 g O_2 will produce $\frac{120}{160} \times 20$ g = **15 g NO.**

Q26. *Using the standard electrode potentials given in the Table 8.1, predict if the reaction between the following is feasible:*

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

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

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

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

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

Ans. (a) $2[Fe^{3+}\,(aq) + e^- \longrightarrow Fe^{2+}\,(aq)]$; $E^0_{red} = 0.77$ V

$2I^-\,(aq) \longrightarrow I_2\,(aq) + 2e^-$; $E^0_{ox} = -0.54$ V

$$2Fe^{3+}\,(aq) + 2I^-\,(aq) \longrightarrow 2Fe^{2+}\,(aq) + I_2\,(aq);$$

$E^0_{reaction} = (0.77 - 0.54)$ V $= +0.23$ V

Since $E^0_{reaction}$ is +ve, the reaction between Fe^{3+} ions and I^- ions is feasible.

(b) $2[Ag^+\,(aq) + e^- \longrightarrow Ag(s)]$; $E^0_{red} = 0.80$ V

$Cu\,(s) \longrightarrow Cu^{2+}\,(aq) + 2e^-$;

$E^0_{ox} = -0.34$ V

$$2Ag^+\,(aq) + Cu\,(s) \longrightarrow 2Ag(s) + Cu^{2+}\,(aq);$$

$E^0_{reaction} = (0.80 - 0.34)$ V

$= +0.46$ V

Since $E^0_{reaction}$ is +ve, the reaction between Ag^+ ions and Cu is feasible.

In a similar way, other reactions can be worked out.

(*c*) $2Fe^{3+} (aq) + Cu(s) \longrightarrow 2Fe^{2+} (aq) + Cu^{2+} (aq);$

$$E^0 = (0.77 - 0.34) V = +0.43 V$$

The reaction is feasible because E^0 is +ve.

(*d*) $Ag (s) + Fe^{3+} (aq) \longrightarrow Ag^+ (aq) + Fe^{2+} (aq);$

$$E^0 = (-0.92 + 0.77) V$$
$$= -0.15 V$$

The reaction is not feasible because E^0 is –ve.

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

$$E^0 = (1.09 - 0.77)V = +0.32 V$$

The reaction is feasible because E^0 is +ve.

Q27. *Predict the products of electrolysis in each of the following:*

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

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

(*iii*) *A dilute solution of H_2SO_4 with platinum electrodes*

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

Ans. (*i*) Ag^+ ions will move towards cathode (–ve electrode) and are discharged:

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

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

(*ii*) At cathode: $Ag^+ (aq) + e^- \rightarrow Ag (s)$

At anode: $2H_2O (l) \rightarrow O_2 (g) + 4H^+ (aq) + 4e^-$

(*iii*) At cathode: $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$

At anode: $2H_2O (l) \rightarrow O_2 (g) + 4H^+ (aq) + 4e^-$

(Note: O_2 is liberated during the electrolysis of salts of oxoanions containing an element in high state of oxidation, such as SO_4^{2-} and NO_3^- ions, *i.e.*, H_2O is more easily oxidised than such an anion.)

(*iv*) At Cathode: $Cu^{2+} (aq) + 2e^- \longrightarrow Cu (s)$

At Anode: $2Cl^- (aq) \longrightarrow Cl_2 (g) + 2e^-$

(When the solution is concentrated)

$$2H_2O (l) \longrightarrow O_2 (g) + 4H^+ (aq) + 4e^-$$

(When the solution is dilute also a small amount of Cl_2 is liberated along with O_2.)

Q28. *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.

Ans. The E^0_{red} potentials of the redox couples of the metals are:
Al^{3+} (aq)/Al (s) = –1.66 V; Cu^{2+} (aq)/Cu (s) = +0.34 V
Fe^{2+} (aq)/Fe (s) = –0.44 V; Mg^{2+} (aq)/Mg(s) = –2.36 V
Zn^{2+} (aq)/Zn (s) = –0.76 V

A metal (A) will displace the other metal (B) from the solution of its ions when E^0_{red} of metal (A) is more negative than metal (B). Thus, the order is Mg, Al, Zn, Fe and Cu.

All metals can displace Cu from Cu^{2+} ions solution.

Q29. *Given the standard electrode potentials,*

$$K^+/K = -2.93\ V,\ Ag^+/Ag = 0.80\ V,$$
$$Hg^{2+}/Hg = 0.79\ V$$
$$Mg^{2+}/Mg = -2.37\ V,\ Cr^{3+}/Cr = -0.74\ V$$

arrange these metals in their increasing order of reducing power.

Ans. More negative E^0_{red} potential value means the metal to be a better reducing agent. Thus, the increasing order of reducing power is:

$$K > Mg > Cr > Hg > Ag$$

Q30. *Depict the galvanic cell in which the reaction*

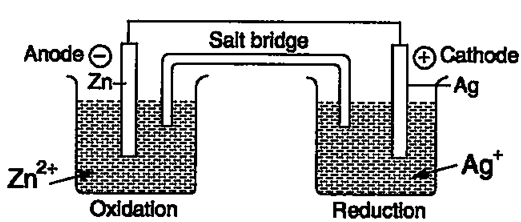
$$Zn\ (s) + 2Ag^+\ (aq) \longrightarrow Zn^{2+}\ (aq) + 2Ag\ (s)$$

takes place.

Further show:

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

Ans. The galvanic cell depicting the redox reaction
$Zn\ (s) + 2Ag^+\ (aq) \longrightarrow Zn^{2+}\ (aq) + 2Ag(s)$



(*i*) Zinc electrode is negatively charged, it is anode because oxidation takes place here.

(*ii*) The ions present in salt bridge are responsible for the flow of current since electricity from one solution to the other flows by the migration of ions through the salt bridge.

(*iii*) At *Anode*, oxidation reaction takes place:

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

At *Cathode*, reduction reaction takes place:

$$Ag^+\,(aq) + e^- \longrightarrow Ag\,(s)$$

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