Unit 8 (REDOX REACTIONS)

Multiple Choice Questions

Single Correct Answer Type

Q1. Which of the following is not an example of redox reaction?

(a) $CuO + H_2 \rightarrow Cu + H_20$

(b) $Fe_20_2 + 3CO \rightarrow 2Fe + 3C0_2$

(c).$2K + F_2 \rightarrow 2KF$

(d) $BaCl_2 + H_2S0_4 \rightarrow BaS0_4 + 2FIC1$

Sol: (d) $BaCl_2 + H_2S0_4$ —> $BaS0_4$ + 2HC1 is not a redox reaction. It is an example of double displacement reactions.

Q2. The more positive the value of E°, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below, find out which of the following is the strongest oxidizing agent. E° values: Fe^{3+}/Fe^{2+} = +0.77; $I_2(g)/I$ = +0.54; Cu^{2+}/Cu = +0.34; Ag^+/Ag = +0.80 V

(a) Fe^{3+}

(b) $I_2(g)$

(c) Cu^{2+}

(d) Ag^+

Sol: (d) Since Ag^+/Ag has highest positive value of E°, therefore, Ag^+ is the strongest oxidizing agent with highest tendency to get reduced

Q3. E° values of some redox complexes are given below. On the basis of these values choose the correct option.

E° values: Br_2/Br^- = +1.90; $Ag^+/Ag(s)$ = +0.80 $Cu^{2+}/Cu(s)$ = +0.34; $I_2(s)/I^-$ = +0.54 V

(a) Cu will reduce Br^-

(b) Cu will reduce Ag

(c) Cu will reduce I^-

(d) Cu will reduce Br_2

Sol: (d) Copper will reduce Br_2, if the E° of the redox reaction, 2Cu + Br_2
$CuBr_2$ is +ve.

$$
\begin{array}{lll}
\text{Now} & Cu \longrightarrow Cu^{2+} + 2e^-; & E^\circ = -0.34\ V \\
& Br_2 + 2e^- \longrightarrow 2Br^-; & E^\circ = +1.90\ V \\
\hline
& Cu + Br_2 \longrightarrow CuBr_2; & E^\circ = +1.56\ V
\end{array}
$$

Since E° of this reaction is +ve, therefore, Cu can reduce Br_2 and hence option (d) is correct.

Q4. Using the standard electrode potential, find out the pair between which redox reaction is not feasible.

E° values: Fe^{3+}/Fe^{2+} = +0.77; I_2/I^- = +0.54;

Cu^{2+}/Cu = +0.34; Ag^+/Ag = +0.80 V

(a) Fe^{3+} and I

(b) Ag^+ and Cu

(c) Fe^{3+} and Cu

(d) Ag and Fe^{3+}

Sol: (d)

(a) $2Fe^{3+} + 2e^- \longrightarrow 2Fe^{2+}; E° = +0.77$ V

$2I^- \longrightarrow I_2 + 2e^-; E° = -0.54$ V (sign of $E°$ is reversed)

$2Fe^{3+} + 2I^- \longrightarrow 2Fe^{2+} + I_2; E°_{cell} = +0.23$ V

This reaction is feasible since $E°_{cell}$ = +ve.

(b) $Cu \longrightarrow Cu^{2+} + 2e^-; E° = -0.34$ V (sign of $E°$ is reversed)

$2Ag^+ + 2e^- \longrightarrow 2Ag; E° = +0.80$ V

$Cu + 2Ag^+ \longrightarrow Cu^{2+} + 2Ag; E° = +0.46$ V

This reaction is feasible since $E°_{cell}$ = +ve.

(c) $2Fe^{3+} + 2e^- \longrightarrow 2Fe^{2+}; E° = +0.77$ V

$Cu \longrightarrow Cu^{2+} + 2e^-; E° = -0.34$ V (sign of $E°$ is reversed)

$2Fe^{3+} + Cu \longrightarrow 2Fe^{2+} + Cu^{2+}; E° = +0.43$ V

This reaction is feasible since $E°_{cell}$ = +ve.

(d) $Ag \longrightarrow Ag^+ + e^-; E° = -0.80$ V (sign of $E°$ is reversed)

$Fe^{3+} + 2e^- \longrightarrow Fe^{2+}; E° = +0.77$ V

$Ag + Fe^{3+} \longrightarrow Ag^+ + Fe^{2+}; E° = -0.03$ V

This reaction is not feasible since $E°_{cell}$ = –ve.

Thus, option (d) is correct.

Q5. Thiosulphate reacts differently with iodine and bromine in the reactions given below:

$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

$S_2O_3^{2-} + 2Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 4Br^- + 10H^+$

Which of the following statements justifies the above dual behaviour of thiosulphate?

(a) Bromine is a stronger oxidant than iodine.

(b) Bromine is a weaker oxidant than iodine.

(c) Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions.

(d) Bromine undergoes oxidation and iodine undergoes reduction in these reactions.

Sol. **(a)** $2\overset{+2}{S_2}\overset{-2}{O_3^{2-}}(aq) + I_2^0(s) \rightarrow \overset{2.5}{S_4}O_6^{2-}(aq) + 2I^-(aq)$

$\overset{+2}{S_2}\overset{-2}{O_3^{2-}}(aq) + 2\overset{0}{Br_2}(l) + 5H_2O(l) \rightarrow 2\overset{+6}{S}\overset{2-}{O_4^{2-}}(aq) + 4Br^-(aq) + 10H^+(aq)$

Bromine being stronger oxidizing agent than I_2, it oxidises S of $S_2O^{2-}{}_3$ to SO_4^{2-} whereas I_2 oxidises it only into $S_4O_6^{2-}$ ion.

Q6. The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following rules is not correct in this respect?
(a) The oxidation number of hydrogen is always +1.
(b) The algebraic sum of all the oxidation numbers in a compound is zero.
(c) An element in the free or the uncombined state bears oxidation number zero.
(d) In all its compounds, the oxidation number of fluorine is -1.
Sol: (a) Oxidation number of hydrogen is -1 in metal hydrides like NaH.

Q7. In which of the following compounds, an element exhibits two different oxidation states.
(a) NH_2OH
(b) NH_4NO_3
(c) N_2H_4
(d) N_3H

Sol: (b) NH_4NO_3 is an ionic compound consisting of NH_4^+ and NO_3^- ion.
The oxidation number of N in two species is different as shown below:
In NH;,
Let the oxidation state of N in NH_4^+ be x.
x + 4x(+l) = +l
x = -3
In NO_3^-
Let the oxidation state of N in NO_3^- be y,
y + 3 x (-2) = -1
y = +5

8. Which of the following arrangements represents increasing oxidation number of the central atom?

(a) $CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-$ (b) $ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-$

(c) $CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}$ (d) $CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-$

Sol. **(a)** Writing the O.N. of Cr, Cl and Mn on each species in the four set of ions, we have,

(a) $\overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^-, \overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-$ (b) $\overset{+5}{Cl}O_3^-, \overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-, \overset{+3}{Cr}O_2^-$

(c) $\overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^-, \overset{+7}{Mn}O_4^-, \overset{+6}{Cr}O_4^{2-}$ (d) $\overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-, \overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^{3-}$

Only in arrangement (a), the O.N. of central atom increases from left to right. Therefore, option (a) is correct.

Q9. The largest oxidation number exhibited by an element depends on its outer electronic configuration. With which of the following outer electronic configurations the element will exhibit largest oxidation number?
(a) $3d^14s^2$
(b) $3d^2\ 4s^2$
(c) $3d^54s^1$
(d) $3d^54s^2$
Sol: (d) Highest O.N. of any transition element = (n – 1)d electrons +ns electrons. Therefore,

larger the number of electrons in the 3d orbitals, higher is the maximum O.N.
(a) $3d^1 4s^2 = 3$; '
(b) $3d^2\ 4s^2 = 3 + 2 = 5$;
(c) $3d^5 4s^1 = 5 + 1 = 6$
(d) $3d^5 4s^2 = 5+2 = 7$

Q10. Identify disproportionation reaction
(a) $CH_4 + 20_2 \rightarrow C0_2 + 2H_20$
(b) $CH_4 + 4C1_2 \rightarrow CC1_4 + 4HCl$
(c) $2F_2 + 20H^- \rightarrow 2F^- + 0F_2 + H_20$
(d) $2N0_2 + 20H^- \rightarrow N0_2 + NO^-{}_3 + H_20$
Sol: (d) Reactions in which the same substance is oxidized as well as reduced are called disproportionation reactions. Writing the O.N. of each element above its symbol in the given reactions,

$$\text{(a)}\quad \overset{-4}{C}\overset{+1}{H_4} + 2\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$$

$$\text{(b)}\quad \overset{-4}{C}\overset{+1}{H_4} + 4\overset{0}{Cl_2} \longrightarrow \overset{+4}{C}\overset{-1}{Cl_4} + 4\overset{+1}{H}\overset{-1}{Cl}$$

$$\text{(c)}\quad 2\overset{0}{F_2} + 2\overset{-2}{O}\overset{+1}{H^-} \longrightarrow 2\overset{-1}{F^-} + \overset{+2}{O}\overset{-1}{F_2} + \overset{+1}{H_2}\overset{-2}{O}$$

$$\text{(d)}\quad 2\overset{+4}{N}\overset{-2}{O_2} + 2\overset{-2}{O}\overset{+1}{H^-} \longrightarrow \overset{+3}{N}\overset{-2}{O_2^-} + \overset{+5}{N}\overset{-2}{O_3^-} + \overset{+1}{H_2}\overset{-2}{O}$$

Thus, in reaction (d), N is both oxidized as well as reduced since the O.N. of N increases from +4 in $N0_2$ to +5 in $N0^-{}_3$ and decreases from +4 in $N0_2$ to +3 in $N0^-{}_2$.

Q11. Which of the following elements does not show disproportionation tendency?
(a) Cl
(b) Br
(c) F
(d) I
Sol: (c) Being the most electronegative element, F can only be reduced and hence it always shows an oxidation number of-1. Further, due to the absence of d-orbitals, it cannot be oxidized and hence it does not show +ve oxidation numbers. In other words, F cannot be simultaneously oxidized as well as reduced and hence does not show disproportionation reactions. Thus, option (c) is correct.

More than One Correct Answer Type

Q12. Which of the following statement(s) is/are not true about the following decomposition reaction?
$2KClO3 \rightarrow 2KC1 + 30_2$
(a) Potassium is undergoing oxidation.
(b) Chlorine is undergoing oxidation.
(c) Oxygen is reduced.
(d) None of the species are undergoing oxidation or reduction.
Sol: (a, b, c, d) Writing the oxidation number of each element above its symbol,

$$2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \longrightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}$$

(a) The O.N. of K does not change, K. undergoes neither reduction nor oxidation. Thus, option

(a) is not correct.
(b) The O.N. of chlorine decreases from +5 in $KCl0_3$ to -l in KCl, hence, Cl undergoes reduction.
(c) Since, O.N. of oxygen increases from -2 in $KC10_3$ to 0 in 0_2, oxygen is oxidized.
(d) This statement is not correct because Cl is undergoing reduction and O is undergoing oxidation.

Q13. Identify the correct statement(s) in relation to the following reaction:
$Zn + 2HCl \rightarrow ZnCl_2 + H_2$
(a) Zinc is acting as an oxidant.
(b) Chlorine is acting as a reductant.
(c) Hydrogen ion is acting as an oxidant.
(d) Zinc is acting as a reductant.

$$\textbf{Sol. (c, d)}\ \overset{0}{Zn} + 2\overset{+1}{H}\overset{-1}{Cl} \longrightarrow \overset{+2}{Zn}\overset{-1}{Cl_2} + \overset{0}{H_2}$$

(a) The O.N. of Zn increases from 0 to +2 (in $ZnCl_2$) and therefore, Zn acts as a reductant and not as an oxidant. Hence, option (a) is not correct.
(b) The O.N. of Cl does not change and therefore, it neither acts as a reductant nor an oxidant. Hence, option (b) is not correct.
(c) The O.N. of H decreases from +1 in H^+ to 0 in H_2. Therefore, H^+ acts an oxidant. This option is correct.
(d) Zinc acts as reductant because its O.N. changes from 0 to +2. This option is correct.

Q14. The exhibition of various oxidation states by an element is also related to the outer orbital electronic configuration of its atom. Atom(s) having which of the following outermost electronic configurations will exhibit more than one oxidation state in its/their compounds.
(a) $3s^1$
(b) $3d^l4s^2$
(c) $3d^24s^2$
(d) $3s^23p^3$

Sol:(b, c, d) Elements which have only s-electron in the valence shell do not show more than one oxidation state. Thus, element with $3s^1$ as outer electronic configuration shows only one oxidation state of +1.
Transition elements, i.e., elements (b, c) having incompletely filled orbitals in the penultimate shell show variable oxidation states. Thus, element with outer electronic configuration as $3d^1$ $4s^2$ shows variable oxidation states of +2 and +3 and the element with outer electronic configuration as $3d^24s^2$ shows variable oxidation states of +2, +3 and +4.
p-Block elements also show variable oxidation states due to a number of reasons such as involvement of J-orbitals and inert pair effect. For example, element (d) with $3s^2$ $3p^3$ as (i.e., P) as the outer electronic configuration shows variable oxidation states of +3 and +5 due to involvement of d-orbitals.

Q15. Identify the correct statements with reference to the given reaction.
$P_4 + 30H^- + 3H_20 \rightarrow PH_3 + 3H_2\ P0^-{}_2$
(a) Phosphorus is undergoing reduction only.
(b) Phosphorus is undergoing oxidation only.
(c) Phosphorus is undergoing oxidation as well as reduction.
(d) Hydrogen is undergoing neither oxidation nor reduction
Sol:

$$\textbf{(c, d)}\ \overset{0}{P_4} + 3\overset{-2}{O}\overset{+1}{H} + 3\overset{+1}{H_2}\overset{-2}{O} \longrightarrow \overset{-3}{P}\overset{+1}{H_3} + 3\overset{+1}{H_2}\overset{-2}{P}\overset{-2}{O_2^-}$$

Because O.N. of P increases from 0 (P_4) to +1 ($H_2P0^-{}_2$) and decreases from 0 (P_4) to -3 (PH_3), therefore, P has undergone both oxidation as well as reduction. So, option (c) is correct. Option (d) is also correct because O.N. of H remains +1 in all the compounds and hence hydrogen is undergoing neither oxidation nor reduction.

Q16. Which of the following electrodes will act as anodes, when connected to Standard Hydrogen Electrode?

(a) $A1^{3-}/A1$; E°= -1.66 V

(b) Fe^{2+}/Fe; E°= -0.44 V

(c) Cu^{2+}/Cu E°=34 V

(d) $F_2(g)/2F^-(aq)$ E°= 2.87 V

Sol:(a, b) The electrodes having negative electrode potentials are stronger reducing agents than H2 gas and therefore, will act as anodes.

Short Answer Type Questions

Q17. The reaction $Cl_2(g) + 20H^-(aq) \rightarrow Cl0^-(aq) + Cl^-(aq) + H_20(l)$ represents the process of bleaching. Identify and name the species that bleaches the substances due to its oxidizing action.

Sol:

$$\overset{0}{Cl_2}(g) + 2\overset{-2}{O}\overset{+1}{H}^-(aq) \rightarrow \overset{+1}{Cl}\overset{-2}{O}^-(aq) + \overset{-1}{Cl}^-(aq) + \overset{+1}{H_2}\ \overset{-2}{O}(l) \text{ (oxidation number)}$$

In this reaction, O.N. of Cl increases from 0 (in Cl_2) to +1 (in ClO^-) and decreases to -1 (in Cl^-). Therefore, Cl_2 is both oxidized to ClO^- and reduced to Cl^-. Since Cl^- ion cannot act as an oxidizing agent (because it cannot decrease its O.N. lower than -1), therefore, Cl_2 bleaches substances due to oxidizing action of hypochlorite, ClO ion.

Q18. $Mn0^{2-}{}_4$ undergoes disproportionation reaction in acidic medium but $Mn0^-{}_4$ does not. Give reason.

Sol: In $Mn0^-{}_4$, Mn is in the highest oxidation state of+7 (i.e.. cannot be oxidized further) and hence it cannot undergo disproportionation.

In contrast, the O.N. of Mn in $Mn0_4^{2-}$ is +6. Therefore, it can increase its O.N. to+7 or decrease its O.N. to some lower value. Thus, $Mn0_4^{2-}$ undergoes disproportionation according to the following reaction.

$$3\overset{+6}{Mn}\overset{-2}{O}{}_4^{2-} + 4H^+ \longrightarrow 2\overset{+7}{Mn}\overset{-2}{O}{}_4^- + \overset{+4}{Mn}\overset{-2}{O}_2 + 2\overset{+1}{H}_2\overset{-2}{O}$$

Here, the O.N. of Mn increases from +6 in $Mn0_4^{2-}$ to +7 in $Mn0^-{}_4$ and decreases to +4 in $Mn0_2$. Thus, $Mn0_4^{2-}$ undergoes disproportionation in acidic medium.

Q19. PbO and $Pb0_2$ react with HC1 according to following chemical equations:

$2PbO + 4HCl \rightarrow 2PbCl_2 + 2H_20$

$Pb0_2 + 4HC1 \rightarrow PbCl_2 + Cl_2 + 2H_20$

Why do these compounds differ in their reactivity?

Sol. (i) $2\overset{+2}{Pb}O + 4HCl \rightarrow 2\overset{+2}{Pb}Cl_2 + 2H_2O$

$$\overset{+4}{Pb}O_2 + 4HCl \rightarrow \overset{+2}{Pb}Cl_2 + \overset{0}{Cl_2} + 2H_2O$$

In reaction (i), O.N. of none of the atoms undergo a change. Therefore, it is not a redox reaction. It is an acid-base reaction, because PbO is a basic oxide which reacts with HCl acid.

The reaction (ii) is a redox reaction in which $Pb0_2$ gets reduced and acts as an oxidizing agent.

Q20. Nitric acid is an oxidizing agent and reacts with PbO but it does not react with $Pb0_2$. Explain why?

Sol: Nitric acid is an oxidizing agent and reacts with PbO to give a simple acid- base reaction without any change in oxidation state. In $Pb0_2$, Pb is in +4 oxidation state and cannot be oxidized further, hence no reaction takes place between $Pb0_2$ and $HN0_3$.

$$2PbO + 4HNO_3 \rightarrow \underset{\text{Salt}}{2Pb(NO_3)_2} \quad + 2H_2O \quad \text{(Acid base reaction)}$$

$$PbO_2 + HNO_3 \longrightarrow \text{No reaction}$$

Q21. Write balanced chemical equations for the following reactions:
(i) Permanganate ion ($Mn0^-{}_4$) reacts with sulphur dioxide gas in acidic medium to produce Mn^{2+} and hydrogensulphate ion. (Balance by ion- electron method)
(ii) Reaction of liquid hydrazine (N_2H_4) with chlorate ion ($C10^-{}_3$) in basic medium produces nitric oxide gas and chloride ion in gaseous state. (Balance by oxidation number method)
(iii) Dichlorine heptaoxide ($C1_20_7$) in gaseous state combines with an aqueous solution of hydrogen peroxide in acidic medium to give chlorite ion ($C10^-{}_2$) and oxygen gas. (Balance by ion-electron method)

Sol: (i) $2Mn0^-{}_4 + 5S0_2 + 2H_20 + H^+ \rightarrow 5HS0^-{}_4 + 2Mn^{2+}$

Balancing by ion-electron method:

$$\overset{+7}{Mn}\overset{-2}{O_4^-} + \overset{+4}{S}\overset{-2}{O_2} \longrightarrow \overset{+2}{Mn^{2+}} + \overset{+1}{H}\overset{+6}{S}\overset{-2}{O_4^-} \quad \text{(Skeletal equation)}$$

Oxidation half: $SO_2 \longrightarrow HSO_4^-$

Reduction half: $MnO_4^- \longrightarrow Mn^{2+}$

Oxidation half: $SO_2 \longrightarrow HSO_4^- + 2e^-$

$$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^- \quad \ldots\text{(i)}$$

(Add $2H_2O$ molecules to balance O atoms)
Reduction half:

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

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

(Add $4H_2O$ molecules to balance O atoms and H atoms)
Add oxidation and reduction half

$$[SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-] \times 5$$

$$[MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O] \times 2$$

$$2MnO_4^- + 5SO_2 + 2H_2O + H^+ \longrightarrow 5HSO_4^- + 2Mn^{2+}$$

(ii) $3N_2H_4 + 4ClO_3^- \rightarrow 6NO + 4Cl^- + 6H_2O$

Balancing by oxidation number method:

O.N. decreases by 6 per Cl atom

$$\overset{-2}{N_2H_4} + \overset{+5}{ClO_3^-} \longrightarrow \overset{+2}{NO} + \overset{-1}{Cl^-}$$

O.N. increases by 4 per N atom

$$6N_2H_4 + 8ClO_3^- \rightarrow 12NO + 8Cl^- + 12H_2O$$

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

Balancing by ion-electron method:

Oxidation half:

$$H_2O_2 \longrightarrow O_2 + 2e^-$$

Adding $2H^+$ on right side to balance H atoms and charge.

$$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$$

Reduction half:

$$Cl_2O_7 + 8e^- \longrightarrow 2ClO_2^-$$

Adding H_2O and H^+ to balance H and O atoms

$$Cl_2O_7 + 8e^- + 6H^+ \longrightarrow 2ClO_2^- + 3H_2O$$

Adding oxidation and reduction half

$$[H_2O_2 \rightarrow O_2 + 2e^- + 2H^+] \times 4$$

$$Cl_2O_7 + 8e^- + 6H^+ \rightarrow 2ClO_2^- + 3H_2O$$

$$Cl_2O_7 + 4H_2O_2 \rightarrow 2ClO_2^- + 3H_2O + 4O_2 + 2H^+$$

Q22. Calculate the oxidation number of phosphorus in the following species.

(a) HPO_3^{2-} and (b) PO_4^{3-}

Sol: (a) Let the oxidation number of P inHPO_3^{2-} be x. So,

$+1 + x + (-6) = -2$ x = +3

(b) Let the oxidation number of P in $PO^3{}_4^-$ be x. So, x + (-8) = -3

x = +5

Q23. Calculate the oxidation number of each sulphur atom in the following compounds:

(a) $Na_2S_2O_3$

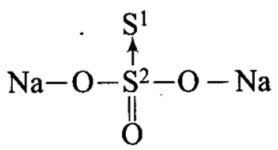
(b) $Na_2S_4O_6$

(c) Na_2SO_3

(d) Na_2SO_4

Sol: (a) $Na_2S_2O_3$

Structure of $Na_2S_2O_3$ is



The oxidation state of S^1 is –2.

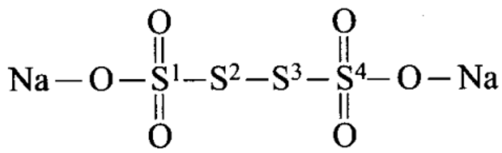
Let oxidation state of S^2 be x.

$$\underset{\text{(For Na)}}{2 \times (+1)} + \underset{\text{(For O)}}{3(-2)} + x + \underset{\text{(For coordinate S)}}{1 \times (-2)} = 0$$

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

Thus, the oxidation states of two S atoms in $Na_2S_2O_3$ are -2 and +6.

(b) $Na_2S_4O_6$



From the left, $S^1 = (-2) + (-2) + (-1) = +5$

$S^2 = 0$

$S^3 = 0$

$S^4 = +5$

(c) Na_2SO_3

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

$x = +4$

(d) Na_2SO_4

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

$x = +6$

Q24. Balance the following equations by the oxidation number method.

(i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

(ii) $I_2 + NO^-_3 \rightarrow NO_2 + IO_3$

(iii) $I_2 + S_2O_3^{2-} \rightarrow I^- + S_4O_6^{2-}$ '

(iv) MnO, + $C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$

Sol. (i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$

$$\overset{+2}{Fe^{2+}} + H^+ + \overset{+6}{Cr_2}\overset{-2}{O_7^{2-}} \longrightarrow \overset{+3}{Cr^{3+}} + \overset{+3}{Fe^{3+}} + H_2O$$

(a) Balance the increase and decrease in O.N.

$$6\overset{+2}{Fe^{2+}} + H^+ + \overset{+6}{Cr_2}\overset{-2}{O_7^{2-}} \longrightarrow 2\overset{+3}{Cr^{3+}} + 6\overset{+3}{Fe^{3+}} + H_2O$$

(b) Balancing H and O atoms by adding H^+ and H_2O molecules

$$6Fe^{2+} + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

(ii) $I_2 + NO_3^- \longrightarrow NO_2 + IO_3^-$

$$I_2 + NO_3^- \longrightarrow NO_2 + IO_3^-$$

$$\overset{0}{I_2} + \overset{+5}{N}\overset{-2}{O_3^-} \longrightarrow \overset{+4}{NO_2} + \overset{+5}{IO_3^-}$$

O.N. decreases by 1

O.N. increases by +5

Total increase in O.N. = 5×2= 10

Total decrease in O.N. = 1

To equalize O.N. multiply NO_3^-, by 10

$I_2 + 10no_3^- \rightarrow 10no_2 + IO_3^-$

Balancing atoms other than O and H

$I_2 + 10no_3^- \rightarrow 10NO_2 + 2\ IO_3^-$

Balancing O and H

I_2 + lO $no_3^- + 8H^+ \rightarrow 10NO_2 + 2\ IO_3^- + 4H_2O$

$$\text{(iii)}\quad \overset{0}{I_2} + \overset{+2}{S_2}\overset{-2}{O_3^{2-}} \longrightarrow \overset{-1}{I^-} + \overset{+2.5}{S_4}O_6^{2-}$$

Total increase in O.N. = $0.5 \times 4 = 2$
Total decrease in O.N. = $1 \times 2 = 2$

To equalize O.N. multiply $S_2O_3^{2-}$ and I^- by 2.

$$I_2 + 2S_2O_3^{2-} \longrightarrow 2I^- + S_4O_6^{2-}$$

$$\text{(iv)}\quad \overset{+4}{Mn}O_2 + \overset{+3}{C_2}O_4^{2-} \longrightarrow \overset{+2}{Mn^{2+}} + \overset{+4}{C}O_2$$

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

To equalize O.N. multiply $C0_2$ by 2. –
$Mn0_2 + C_20^{2-}{}_4 \rightarrow Mn^{2+} + 2CO_2$
Balance H and O by adding $2H_20$ on right side, and $4H^+$ on left side of equation.
$Mn0_2 + C_20^{2-}{}_4 + 4H^+ \rightarrow Mn^{2+} + 2CO_2 + 2H_20$

Q25. Identify the redox reactions out of the following reactions and identify the oxidizing and reducing agents in them

(i) $3HCl(aq) + HNO_3(aq) \rightarrow Cl_2(g) + NOCl(g) + 2H_2O(l)$
(ii) $HgCl_2(aq) + 2KI(aq) \rightarrow HgI_2(s) + 2KCl(aq)$
(iii) $Fe_2O_3(s) + 3CO(g) \xrightarrow{\Delta} 2Fe(s) + 3CO_2(g)$
(iv) $PCl_3(l) + 3H_2O(l) \rightarrow 3HCl(aq) + H_3PO_3(aq)$
(v) $4NH_3 + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(g)$

Sol: (i) Writing the O.N. of on each atom,

$$3\overset{+1-1}{HCl}(aq) + \overset{+1}{H}\,\overset{+5}{N}\,\overset{-2}{O_3}(aq) \longrightarrow \overset{0}{Cl_2}(g) + \overset{+3}{N}\,\overset{-2}{O}\,\overset{-1}{Cl}(g) + 2\overset{+1}{H_2}\overset{-2}{O}(l)$$

Here, O.N. of Cl increases from -1 (in HCl) to 0 (in Cl_2). Therefore, Cl^- is oxidized and hence HCl acts as a reducing agent.
The O.N. of N decreases from +5 (in $HN0_3$) to +3 (in NOC1) and therefore, $HN0_3$ acts as an oxidizing agent.
Thus, reaction (i) is a redox reaction.

$$\text{(ii)}\quad \overset{+2-1}{HgCl_2}(aq) + 2\overset{+1-1}{K\,I}(aq) \longrightarrow \overset{+2}{Hg}\,\overset{-1}{I_2}(aq) + \overset{+1-1}{KCl}(aq)$$

Here O.N. of none of the atoms undergo a change and therefore, this is not a redox reaction.

$$\text{(iii)}\quad \overset{+3-2}{Fe_2O_3}(s) + 3\overset{+2\,-2}{CO}(g) \xrightarrow{\Delta} 2\overset{0}{Fe}(s) + 3\overset{+4}{C}\overset{-2}{O_2}(g)$$

Here, O.N. of Fe decreases from +3 (in Fe_20_3) to 0 (in Fe) and therefore, Fe_20_3 acts as an oxidizing agent.
O.N. of C increases from +2 (in CO) to +4 (in $C0_2$) and therefore, CO acts as a reducing agent.
Thus, this is a redox reaction.

$$\text{(iv)}\quad \overset{+3}{P}\overset{-1}{Cl_3}(l) + 3\overset{+1}{H_2}\overset{-2}{O}(l) \longrightarrow 3\overset{+1}{H}\overset{-1}{Cl}(aq) + \overset{+1}{H_3}\overset{+3}{P}\overset{-2}{O_2}(aq)$$

Here O.N. of none of the atoms undergo a change and therefore, it is not a redox reaction.

$$\text{(v)}\quad 4\overset{-3}{N}\overset{+1}{H_3}(g) + 3\overset{0}{O_2}(g) \longrightarrow 2\overset{0}{N_2}(g) + 6\overset{+1}{H_2}\overset{-2}{O}(g)$$

Here. O.N. of N increases from -3 (in NH_3) to 0 in (N_2) and therefore, NH_3 acts as a reducing agent. O.N. of O decreases from 0 (in 0_2) to -2 (in H,0) and therefore, 0_2 acts as an oxidizing agent.

Thus, reaction (v) is a redox reaction.

26. Balance the following ionic equations.

(i) $Cr_2O_7^{2-} + H^+ + I^- \rightarrow Cr^{3+} + I_2 + H_2O$

(ii) $Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

(iii) $MnO_4^- + SO_3^{2-} + H^+ \rightarrow Mn^{2+} + SO_4^{2-} + H_2O$

(iv) $MnO_4^- + H^+ + Br^- \rightarrow Mn^{2+} + Br_2 + H_2O$

Reduction

Sol. (i) $\overset{+6}{Cr_2}O_7^{2-} + H^+ + \overset{-1}{I^-} \longrightarrow \overset{+3}{Cr^{+3}} + \overset{0}{I_2} + H_2O$

Oxidation

Dividing the equation into two half reactions:

Oxidation half reaction: $I^- \rightarrow I_2$

Reduction half reaction: $Cr_20_7^{2-} \rightarrow Cr^{3+}$

Balancing oxidation and reduction half reactions separately as:

Oxidation half reaction

$$I^- \longrightarrow I_2$$

$$2I^- \longrightarrow I_2$$

$$2I^- \longrightarrow I_2 + 2e^- \qquad \ldots\text{(i)}$$

Reduction half reaction

$$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$$

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$$

$$Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+}$$

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O \qquad \ldots\text{(ii)}$$

(acidic medium)

To balance the electrons, multiply eq. (i) by 3 and add to eq. (ii)

$$Cr_2O_7^{2-} + 14H^+ + 6I^- \longrightarrow 2Cr^{3+} + 3I_2 + 7H_2O$$

(ii) Step-1: Separate the equation into two half reactions.

The oxidation number of various atoms are shown below:

$$\overset{+6\ -2}{Cr_2O_7^{2-}} + \overset{+2}{Fe^{2+}} + \overset{+1}{H^+} \rightarrow \overset{+3}{Cr^{3+}} + \overset{+3}{Fe^{3+}} + \overset{+1\ -2}{H_2O}$$

In this case, chromium undergoes reduction, oxidation number decreases from +6 (in $Cr_2O_7^{2-}$) to +3 (in Cr^{3+})

Fe^{2+} (O.N. = +2) changes to Fe^{3+} (O.N. = +3). The species undergoing oxidation and reduction are:

Oxidation: $Fe^{2+} \longrightarrow Fe^{3+}$

Reduction: $Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

Step–2: Balance each half reaction separately as:

(a) $Fe^{2+} \longrightarrow Fe^{3+}$

(i) Balance all atoms other than H and O. This step is not needed, because, it is already balanced.

(ii) The oxidation number on left is +2 and on right is +3. To account for the difference, the electron is added to the right as:

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

(iii) Charge is already balanced.

(iv) No need to add H or O.

The balanced half equation is:

$Fe^{2+} \rightarrow Fe^{3+} + e^-$ ...(i)

Consider the second half equation

$Cr_20_7^{2-} \rightarrow Cr^{3+}$

(i) Balance the atoms other than H and O.

$Cr_20_7^{2-} \rightarrow 2Cr^{3+}$

(ii) The oxidation number of chromium on the left is +6 and on the right is +3. Each chromium atom must gain three electrons. Since there are two Cr atoms, add $6e^-$ on the left.

$Cr_20_7^{2-} + 6e^- \rightarrow 2Cr^{3+}$

(iii) Since the reaction takes place in acidic medium add $14H^+$ on the left to equate the net charge on both sides.

$Cr_20_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+}$

(iv) To balance Fl atoms, add $7H_20$ molecules on the right.

$Cr_20_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_20$ .. .(ii)

This is the balanced half equation.

Step-3: Now add up the two half equations. Multiply eq. (i) by 6 so that electrons are balanced.

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-] \times 6$$

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$$

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \longrightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

The balanced equation is:

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \longrightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

Reduction

(iii) $\overset{+7}{Mn}O_4^- + \overset{+4}{S}O_3^{2-} \longrightarrow \overset{+2}{Mn}^{2+} + \overset{+6}{S}O_4^{2-} + H_2O$

Oxidation

Dividing the equation into two half reactions:

Oxidation half reaction: $SO_3^{2-} \longrightarrow SO_4^{2-}$

Reduction half reaction: $MnO_4^- \longrightarrow Mn^{2+}$

Balancing oxidation and reduction half reactions separately as:

Oxidation half reaction

$$SO_3^{2-} \longrightarrow SO_4^{2-}$$

$$SO_3^{2-} \longrightarrow SO_4^{2-} + 2e^-$$

Since the reaction occurs in acidic medium,

$$SO_3^{2-} \longrightarrow SO_4^{2-} + 2e^- + 2H^+$$

$$SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2H^+ + 2e^- \qquad ...(i)$$

Reduction half reaction

$$MnO_4^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

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

To balance the electrons, multiply eq. (i) by 5 and eq. (ii) by 2 and add

$$2MnO_4^- + 5SO_3^{2-} + 6H^+ \longrightarrow 2Mn^{2+} + 5SO_4^{2-} + 3H_2O$$

Reduction

(iv) $\overset{+7}{Mn}O_4 + H^+ + \overset{-1}{Br^-} \longrightarrow \overset{+2}{Mn}^{2+} + \overset{0}{Br_2} + H_2O$

Oxidation

Dividing the equation into two half reactions:

Oxidation half reaction: $Br^- \longrightarrow Br_2$

Reduction half reaction: $MnO_4^- \longrightarrow Mn^{2+}$

Balancing oxidation and reduction half reactions separately as:

Oxidation half reaction

$$Br^- \longrightarrow Br_2$$

$$2Br^- \longrightarrow Br_2$$

$$2Br^- \longrightarrow Br_2 + 2e^- \qquad ...(i)$$

Reduction half reaction

$$MnO_4^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+}$$

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

To balance the electrons, multiply eq. (i) by 5 and eq. (ii) by 2 and add

$$2MnO_4^- + 10Br^- + 16H^+ \longrightarrow 2Mn^{2+} + 5Br_2 + 8H_2O$$