Date: 16/10/2021

Q1. Can the reaction:

$$Cr_2O_7^{2-} + H_2O \longrightarrow 2CrO_4^{2-} + 2H^+$$

be regarded as a redox reaction?

Q2. Name a non-metal that does not show disproportionation reaction.

Q3. Name five non-metals that can show disproportionation reaction.

Q4. Fluorine (F_2) does not show disproportionation. Give reason.

Q5. The oxoanion ClO_4^- does not show disproportionation reaction. Give reason.

Q6. Show that $H_2O\,(s) + F_2\,(g) \rightarrow HF\,(g) + HOF\,(g)$ is a redox reaction.

Q7. Find the change in the oxidation number of S in H_2O and SO_2 in the following reaction.
$2H_2S\,(s) + SO_2\,(g) \rightarrow 3S\,(s) + H_2O\,(g)$

Q8. Try all possible approaches to justify that the following reactions are redox reactions.

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

Q9. Try all possible approaches to justify that the following reactions are redox reactions.

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

Q10. Try all possible approaches to justify that the following reactions are redox reactions.

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

Q11. Try all possible approaches to justify that the following reactions are redox reactions.

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

Q12. At what concentration of $Cu^{2+}(aq)$ will its electrode potential become equal to its standard electrode potential?

Q13. An iron rod is immersed in a solution containing $NiSO_4$ and $ZnSO_4$. When the concntration of each salt is 1 M, predict giving reasons which of the following reactions is likely to proceed?

(a) Iron reduces Zn^{2+} ions (b) Iron reduces Ni^{2+} ions

Given : $E^\circ_{(Zn^{2+}|Ni)} = -0.25$ V

Q14. Nitric acid is an oxidising agent and reacts with PbO but it does not react with PbO_2. Explain why?

Q15. Identify the oxidising and reducing agent in the following reactions.

$$Fe^{2+} + 2H^+ + NO_3^- \longrightarrow Fe^{3+} + NO_2 + H_2O$$

Q16. Calculate the oxidation number of Fe in

(a) Fe_3O_4 (b) $Fe_4[Fe(CN)_6]_3$

Q17. The electrode reduction potentials of four metallic elements A, B, C and D are respectively + 0.79, –0.74, 1.08 and –0.31 V. Arrange these in order of decreasing electropositive character.

Q18. Assign oxidation number to the underlined elements in each of the following species:
(a) $Ca\underline{O}_2$ (b) $NaB\underline{H}_4$ (c) $H_2\underline{S}_2O_7$ (d) $KAl(\underline{S}O_4)_2 12H_2O$

Q19. Write formulas for the following compounds:
(a) Mercury (II) chloride (b) Nickel (II) sulphate (c) Tin (IV) oxide
(d) Thallium (I) sulhate (e) Iron (III) sulphate (f) Chromium (III) oxide

Q20. Assign oxidation number to the underlined elements in each of the following species:
(a) $NaH_2\underline{P}O_4$ (b) $NaH\underline{S}O_4$ (c) $H_4\underline{P}_2O_7$ (d) $K_2\underline{Mn}O_4$

Q21. The compound AgF_2 is unstable. However, if formed, the compund acts as a very strong oxidising agent. Why?

Q22. 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) $HCHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \longrightarrow 2Ag(s) + HCOO^-(aq) + 4NH_3(aq) + 2H_2O(l)$

Q23. Formation of glucose from carbon dioxide and water during photosynthesis takes place as:

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

Give reason to represent this reaction as follows:

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

Q24. Write four informations about the reaction:

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

Q25. 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?

Q26. Using reaction $Pb_3O_4 + 8HCl \rightarrow 3PbCl + 3PbCl_2 + Cl_2 + 4H_2O$ justify that Pb_3O_4 is stoichiometric mixture of 2 moles of PbO and one mole of PbO_2.

Q27. The oxoanions ClO^-, ClO_2^- and ClO_3^- Show disproprtionation reaction. Give reason.

Q28. With reference to the following reactions draw inference about the behaviour of Ag^+ and Cu^{2+}:
(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) $C_6H_5CHO(l) + 2[Ag(NH_3)]^+(aq) + 3OH^-(aq) \longrightarrow C_6H_5COO^-(aq) + 2Ag(s) + 4NH_3\,(aq) + 2H_2O(l)$
(d) $C_6H_5CHO(l) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow$ No change observed

Q29. Justify that the reaction $Pb_3O_4 + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + PbO_2 + 2H_2O$ is essentially acid base reacition though Pb_3O_4 is a stoichiometric mixture of 2 moles of PbO and one mole of PbO_2 which is an oxidising agent.

Q30. Nitric acid acts only as an oxidising agent while nitrous acid acts both as an oxidising as well as reducing agent. Explain.

Q31. Split the following redox reactions in the oxidation and reduction half reactions :
(a) $2K\,(s) + Cl_2\,(g) \longrightarrow 2KCl\,(s)$
(b) $2Al(s) + 3Cu^{2+}(aq) \longrightarrow 2Al^{3+}\,(aq) + 3Cu(s)$

Q32. Arrange the following molecules in the decreasing order of oxidation state (+ ve to – ve) of nitrogen : NO_2, HN_3, NO_2^-, N_2H_4.

Q33. Which one of the following is not a redox reaction? Justify your answer.

(a) $2NO_2(g) + 4OH^-(aq) \rightarrow NO_2^-(aq) + NO_3^-(aq) + H_2O(l)$

(b) $Na_2SO_4(g) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl$

Q34. Calculate the oxidation number of the underlined element in

(a) $\underline{V}O_2^+$ (b) $\underline{U}O_2^{2+}$ (c) Ba_2XeO_6 (d) $K_2\underline{S}$

Q35. Can the following reaction : $Cr_2O_7^{2-} + H_2O \rightleftharpoons 2CrO_4^{2-} + 2H^+$ be regarded as a redox reaction?

Q36. Which of the following reactions are feasible?

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

(b) $I_2(s) + 2Ag^+(aq) \longrightarrow 2I^-(aq) + Br_2$

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

(d) $2Ag + 2H^+(aq) \longrightarrow H_2 + 2Ag^+(aq)$

Q37. Predict the maximum and minimum oxidation states for (i) Cl (ii) Ti

Q38. Which of the following equations represent oxidation reduction reaction? Identify each oxidising agent and each reducing agent.

(a) $KOH + H_2O_2 \longrightarrow KOH_2 + H_2O$

(b) $Cr_2O_7^{2-} + 2OH^- \longrightarrow 2CrO_4^{2-} + H_2O$

(c) $K + O_2 \longrightarrow KO_2$

(d) $Ca(HCO_3)_2 \longrightarrow CaCO_3 + CO_2 + H_2O$

Q39. What is the oxidation number of N in HNO_4?

Q40. 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.

Q41. Calculate the oxidation number of nickel in $Ni(CO)_4$, iron in $Fe(CO)_5$ and carbon in CH_2O.

Q42. What are the oxidation number of the underlined elements in each of the following and how do you rationalize your results? (a) $\underline{C}H_3\underline{C}H_2OH$; (b) $\underline{C}H_3\underline{C}OOH$.

Q43. What are the oxidation number of the underlined elements in each of the following and how do you rationalize your results? (a) $\underline{KI}_3$; (b) $\underline{Fe}_3O_4$.

Q44. Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_4, CrO_5 and NO_3^-. Suggest structures of these three compounds. Count for the fallacy.

Q45. Fluorine reacts with ice and results in the change:

$$\overset{+1\ -2}{H_2O}(s) + \overset{0}{F_2}(g) \longrightarrow \overset{+\ -1}{HF} + \overset{+1\ 0\ -1}{HOF}(g)$$

Justify that this reaction is a redox reaction. If oxygen of HOF disproportionates at room temperature, then what reaction is possible?

Q46. While sulphur dioxide and hydrogen peroxide can act as an oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

Q47. Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions.

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

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

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

Q48. 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 the reaction for this redox change taking place in water.

Q49. Consider the elements: Cs, Ne, I, F

(a) Identify the element that exhibits –ve oxidation state.

(b) Identify the element that exhibits +ve oxidation state.

(c) Identify the element that exhibits both +ve and –ve oxidation state.

(d) Identify the element which neither exhibits –ve nor +ve oxidation state.

Q50. How does Cu_2O act as both oxidant and reductant? Explain with proper reactions showing the change of oxidation number in each case.

Q51. When 10 mL of a solution of Fe^{2+} ions was titrated in the presence of dil H_2SO_4 using diphenylamine indicator, 15 mL of 0.02 M solution of $K_2Cr_2O_7$ was used to get the end point. Calculate the molarity of the solution containing Fe^{2+} ions.

Q52. Giving reason select the type of reactions as disproportionation, combinaton, decompositon, displacement from the following list.

(a) $2\,Pb(NO_3)_2(s) \rightarrow 2\,PbO(s) + 4\,NO_2\,(g) + \frac{1}{2}\,O_2(g)$

(b) $2\,NO_2\,(g) + 4\,OH^-\,(aq) \rightarrow NO_2^-\,(aq) + NO_3^-\,(aq) + H_2O(l)$

(c) $N_2(g) + O_2(g) \rightarrow 2NO\,(g)$

Date: 16/10/2021

S1. In this reaction,

Oxidation number of Cr in $Cr_2O_7^{2-}$ = + 6

Oxidation number of Cr in CrO_4^{2-} = + 6

Since during this reaction, the oxidation number of Cr has neither increased nor decreased. therefore, the above reaction cannot be regarded as a redox reaction.

S2. Fluorine (F_2) does not show disproportionation reaction.

S3. Chlorine (Cl_2), bromine (Br_2), iodine (I_2), phosphorus (P_4), sulphur (S_8)

S4. In F_2 the oxidation number of fluorine is 0 which is its highest value. Therefore, F_2 doest not show disproportionation reaction.

S5. In the oxoanion ClO_4^- the oxidation number of chlorine is +7 which is its highest value. Therefore, the oxanion ClO_4^- doest not show disproportionationreaction.

Justification: In order to show disproportionation reaction the main atom of the species should not be in its highest or lowest oxidation state but it should have intermediate oxidation number.

S6.

$$\underset{+1\ -2}{H_2O\,(s)} + \underset{0}{F_2\,(g)} \longrightarrow \underset{+1\ -1}{HF(g)} + \underset{0.5\ -1}{HOF\,(g)}$$

o.n. +1 –2 0 +1 – 1 0.5 – 1

There is change in the oxidation number of each atom of the above reaction. Therefore, it is a redox reaction.

S7. We shall write the oxidation number of S atom in the reactants and products below each symbol.

$$\underset{+1\ -2}{2H_2S\,(s)} + \underset{+4\ -2}{SO_2\,(g)} \longrightarrow \underset{0}{3S\,(s)} + \underset{+1\ -2}{H_2O\,(g)}$$

o.n. +1 – 2 +4 – 2 0 +1 – 2

- In H_2S the oxidation number of S atom changes from –2 to 0.
- In SO_2 the oxidation number of S atom changes from +4 to 0.

S8. The chemical equation along with oxidation number of each element will be written as:

$$\overset{0}{2K}(s) + \overset{0}{F_2}(g) \longrightarrow 2K^+F^-(s)$$

Each K atom has 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 *redox reaction*.

S9. The chemical equation along with oxidation number of each element is:

$$\overset{+3\ -1}{4BrCl_3(g)} + \overset{+1\,+3\,-1}{3LiAlH_4(s)} \longrightarrow \overset{-3\ \ +1}{2B_2H_6(g)} + \overset{+1\ -1}{2LiCl(s)} + \overset{+3\ -1}{3AlCl_3(s)}$$

Here, O.N. of B decreases from +3 in BCl_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 $BrCl_3$ but is removed from $LiAlH_4$, therefore, BCl_3 is reduced while $LiAlH_4$ is oxidised. Thus it is a *redox reaction*.

S10. The chemical equation of the reaction is:

$$\overset{+3}{Fe_2}\overset{-2}{O_3}(s) + 3\overset{+2}{C}O(g) \longrightarrow 2\overset{0}{Fe}(s) + 3\overset{-4}{CO_2}(g)$$

Here, O.N. of Fe decreases from + 3 in Fe_2O_3 to 0 in Fe wile that of C increases from +2 in CO to +4 in CO_2. Further, oxygen is removed from Fe_2O_3 and added in CO, therefore, FeO_3 is reduced while CO is oxidised. Thus this is a *redox reaction*.

S11. The chemical equation with oxidation number of each element shown on its symbol is written as:

$$\overset{+2}{Cu}\,\overset{-2}{O}(s) + \overset{0}{H_2}(g) \longrightarrow \overset{0}{Cu}(s) + \overset{+1}{H_2}\,\overset{-2}{O}(g)$$

Here, O is removed from CuO, therefore, it is reduced to Cu while O is added to H_2 to form H_2O, therefore hydrogen 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*.

S12. At 1 M concentration.

S13. (a) The reduction potential of iron is more than that of zinc. Therefore, iron will be reduced. In other words, Zn^{2+} will not be reduced by iron.

(b) The reduction potential of Ni^{2+} is more than that of iron. Therefore, Ni^{2+} will be reduced by iron.

S14. PbO is a basic oxide and reacts with. NHO_3 and simple acid base reaction take place.

$$2PbO + 4HNO_3 \longrightarrow 2Pb(NO_3)_2 + 2H_2O$$

But in PbO_2, lead is in +4 oxidation state and cannot be further oxidised. Therefore, it does not react with HNO_3.

S15.

$$\overset{+2}{Fe^{2+}} + 2H^+ + NO_3^- \longrightarrow \overset{3+}{Fe^{3+}} + \overset{+4}{NO_2} + H_2O$$

Fe^{2+} is oxidised to Fe^{3+} by NO_3^- ion and therefore, NO_3^- is an oxidising agent. NO_3^- is reduced from N (+5) to N(+4) in NO_2 by Fe^{2+} and therefore, Fe^{2+} is a reducing agent.

S16. (a) Fe_3O_4 $\quad$ $3x - 8 = 0$ or $x = 8/3$ (fractional)

Actually the composition of $FeSO_4$ is $FeO.Fe_2O_3$

The O.N. of Fe in FeO is +2 and in Fe_2O_3, it is +3.

On the average O.N. comes out to be 8/3.

(b) $Fe_4[Fe(CN)_6]_3$. In this case the O.N. of Fe in the bracket is +2 and that of Fe outside the brackets is +3. $Fe_4^{+3}[Fe^{+2}(CN)_6]_3$

S17. Higher the electrode reduction potential, lower is its tendency to lose electrons and therefore, lower is the electrpositive character of the metal. So, the metals can be arranged as : B, D, A, C.

S18. (a) The oxidation number of Ca is +2. Thus, the sum of oxidation numbers in $\overset{+2}{Ca}\overset{x}{\underline{O}}_2$ wil be

$$+2 + 2x = 0 \quad \text{or} \quad x = -1$$

Thus, oxidation number of oxygen in CaO_2 = **– 1**.

(b) In $NaBH_4$, H is present as hydride ion. There, its oxidation number is – 1. Thus, for $\overset{+1}{Na}\overset{x}{\underline{B}}\overset{-2}{H_4}$, the sum of oxidation numbers is $1(+1) + x + 4(-1) = 0$ or $x + 3$.

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

(c) The sum of oxidation numbers in, $\overset{+1}{Na_2}\overset{x}{\underline{S}_2}\overset{-2}{O_7}$, will be given as

$$2(+1) + 2(x) + 7(-2) = 0 \quad \text{or} \quad x = +6.$$

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

(d) The sum of oxidation numbers in, $\overset{+1}{K}\overset{+3}{Al}(\overset{x}{\underline{S}}\overset{-2}{O_4})_2 12(\overset{+1}{H_2}\overset{-2}{O})$, is

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

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 can be ignored while computing the oxidation number of S. Accordingly,

$$+1 + 3 + 2x - 16 = 0 \quad \text{or} \quad x = +6$$

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

S19. (a) $Hg(II)Cl_2$ (b) $Ni(II)SO_4$ (c) $Sn(IV)O_2$

(d) $Th_2(I)SO_4$ (e) $Fe_2(III)(SO_4)_3$ (f) $Cr_2(II)O_3$

S20. (a) Let the oxidation number of P be x. The oxidation number of each atom is shown above its symbol,

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

Sum of oxidation numbers of various atoms in NaH_2PO_4 is zero, hence

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

or $\quad x - 5 = 0 \quad$ or $\quad x = +5$

Thus, the oxidation number of P in NaH_2PO_4 = **+ 5**.

(b) In the similar way as discussed, the sum of oxidation numbers in $\overset{+1}{Na}\overset{+1}{H}\overset{x}{\underline{S}}\overset{-2}{O_4}$ will be,

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

Thus, the oxidation number of S in $NaHSO_4$ = **+ 6**.

(c) The sum of oxidation numbers in, $\overset{+1}{H_4}\overset{x}{\underline{P}_2}\overset{-2}{O_7}$, will be

$$4(+1) + 2(x) + 7(-2) = 0 \quad \text{or} \quad x = +5$$

Thus, the oxidation number of P in $H_4P_2O_7$ = **+5**.

(d) The sum of of oxidation numbers in $\overset{+1}{K_2}\overset{x}{\underline{Mn}}\overset{-2}{O_7}$, will be

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

Thus, the oxidation number of Mn in K_2MnO_4 = **+6**.

S21. In AgF_2, oxidation state of Ag is +2 which is very very unstable. Since Ag can exist in a stable state of +1 therefore, it quickly accepts an electron to form the more stable +1 oxidation state.

$$Ag^{2+} + e^- \longrightarrow Ag^+$$

Therefore, AgF_2, if formed, will act as a strong oxidising agent.

S22. The list of substance oxidised, reduced, etc., for the given reactions are tabulated below:

	Substance oxidised	*Substance reduced*	*Oxidising agent*	*Reducing agent*
(a)	$C_6H_6O_2$(aq)	AgBr(s)	AgBr(s)	$C_6H_6O_2$(aq)
(b)	HCHO(*l*)	$[Ag(NH_3)_2]^+$	$[Ag(NH_3)_2]^+$	HCHO(*l*)

S23. During Photosynthesis in the presence of sunlight and chlorophyll water is oxidised to give O_2. Now 6 molecules of O_2 are produced by the oxidation of 12 molecules of H_2O. In the other part of the reaction CO_2 reacts with hydrogen given by water to form glucose and water. Therefore,the reaction is represented as

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

S24. (a) It is a disproportionation reaction.

(b) Cyanogen $(CN)_2$ gets simultaneously reduced to CN^- ion as well as oxidised to cyanate ion.

(c) O.N. of N in $(CN)_2$ is –3 while that in CN^- is –2 and in CNO^- is –5.

(d) The reaction occurs in basic medium.

S25. The average O.N. of S in $S_2O_3^{2-}$ is +2 while in $S_4O_6^{2-}$ it is +2.5. The O.N, of S in SO_4^{2-} is +6. Since Br_2 is a stronger oxidising agent that I_2, it oxidies S of $S_2O_3^{2-}$ to a higher oxidation state of +6 and hence forms SO_4^{2-} ion. I_2, however, being a weaker oxidising agent oxidies S of $S_2O_3^{2-}$ ion to a lower oxidation of +2.5 in $S_4O_6^{2-}$ ion. It is because of this reason that thiosulphate reacts differently with Br_2 and I_2.

S26. (a) PbO is a basic oxide and it reacts with hydrochloric acid as

$$2\,PbO + 4HCl \longrightarrow 2PbCl_2 + Cl_2$$

(b) PbO_2 is an oxidising agent, it oxidises Cl^- ion of HCl to Cl_2 whereas Pb is reduced from + 4 oxidation state to + 2 state

o.n.	+4		−1		+2		0	
	PbO_2	+	4HCl	$\longrightarrow$	$PbCl_2$	+	$2H_2O$	... Redox reaction

On adding the above two ractions we shall get

$(2\,PbO + PbO_2) + 8\,HCl \rightarrow 3\,PbCl_2 + Cl_2 + 4\,H_2O$ ... Overall reaction

Thus it is proved that Pb_3O_4 is a stoichiometric mixture of 2 moles of PbO and 1 mole of PbO_2. That is, $Pb_3O_4 \equiv (PbO + PbO_2)$.

S27. The oxoanions ClO^-, ClO^-_2 and ClO^-_3 show disproprtionation reactions because the chlorine atom in each one of them is in its intermediate oxidation state.

(a) o.n.Cl	+1		−1		+5	
	$3ClO^-$	$\rightarrow$	$2Cl^-$	+	ClO_3^-	
(b) o.n.Cl	+3		− 1		+ 5	
	$3ClO_2^-$		Cl^-		ClO_3^-	
(c) o.n.Cl	+ 5		−1		+7	
	$4ClO_3^-$	$\rightarrow$	Cl^-	+	$3ClO_4^-$	

S28. In reactions (a) and (b) both Ag and Cu^{2-} act as oxidising agents.

From reactions (c) and (*d*) it is inferred that

(a) Ag^+ is better oxidising agent than Cu^{2+}

(b) Cu^{2+} has lesser tendency to get reduced as compared to Ag^+

(c) Cu is more reactive than Ag.

S29. In $Pb_3O_4 \equiv (2\,PbO + PbO_2)$ it is true that PbO_2 is an oxidising agent. But it is ineffective in the presence of NHO_3 which itself is a strong oxidising agent. Thus PbO_2 cannot oxidise HNO_3.

$2\,PbO + 4\,HNO_3 \rightarrow 2\,Pb(NO_3)_2 + 2\,H_2O$

$\rightarrow$ $Pb_3O_4 \equiv (2\,PbO + PbO_2) + 4\,HNO_3 \rightarrow 2\,Pb(NO_3)_2 + PbO_2 + 2\,H_2O$.

S30. (a) Oxidation number of N in HNO_3 is + 5.

Maximum oxidation number of N is +5 because it has five electrons the valence shell ($2s^2\ 2p^3$).

Minimum oxidation number of N is – 3 because it has accept 3 more electrons to get noble configuration.

Since O.N. of N in HNO_3 is maximum therefore, it can only decrease. Thus, HNO_3 can act as an oxidising agent.

(b) HNO_2. Oxidation number of N = + 3

Maximum O.N. of N = + 5

Minimum O.N. of N = – 3

Therefore, the O.N. of N can increase by losing electrons or can decrease by accepting electrons. Thus, HNO_2 can act both as an oxidising as well as a reducing agent.

S31. (a) Oxidation half reaction : $2K(s) \longrightarrow 2K^+(aq) + 2e^-$

Reduction half reaction : $Cl_2(g) + 2e^- \longrightarrow 2Cl^-$

(b) Oxidation half reaction : $2Al(s) \longrightarrow 2Al^{3+}(aq) + 3e^-$

Reduction half reaction : $3Cu^{2+} + 6e^- \longrightarrow 3Cu(s)$

S32. Oxidation states of N are :

$NO_2 = +4$, $NH_3 = -3$, $HN_3 = -\frac{1}{3}$

$NO_2^- = +3$, $N_2H_4 = -2$

Decreasing order of oxidation state:

$$NO_2 > NO_2 > NH_3 > N_2H_4 > NH_3$$

S33. Reaction (b) does not represent a true redox reaction. It is a double displacement reaction.

Justification: There is no change in the oxidation number of any atom involved in the given reaction

o.n. $\quad +1\,+6\,-2 \quad + \quad 2\,-1 \qquad +2\,+6\,-2 \quad +1\,-1$

$Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl$

S34. The O.N. of underlined atom is represented by x

(a)	$\underline{V}O_2^+$	$x + 2(-2) = +1$	or	$x = +5$
(b)	$\underline{U}O_2^{2+}$	$x + 2(-2) = +2$	or	$x = +6$
(c)	$BaXeO_6$	$+2(2) + x + 6(-2) = 0$	or	$x = +5$
(d)	K_2S	$1 \times 2 + x = 0$	or	$x = -2$

S35. In this reaction, oxidation number of Cr in $Cr_2O_7^{2-}$ is + 6 and oxidation number of Cr in CrO_4^{2-} is + 6. Since during the reaction, the oxidation number of Cr has neither decreased nor increased, therefore, the above reaction is not a redox reaction.

S36. (a) and (c) are feasible

S37. Maximum oxidation state for Cl = +7

Minimum oxidation state for Cl = –1

Maximum oxidation state for Ti = + 4

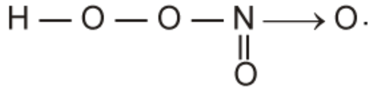
Minimum oxidation state for Ti = 0

For metals minimum O.N. is zero but for non-metals, it is group number minus 8.

S38. (a) *a*, *b*, and *d* are not redox reactions.

c is a redox reaction. The oxidising agent is O_2 and reducing agent is K.

S39. By conventional method, O.N. of N is : +1 + x + 4(– 2) = 0 or x = +7.

However, N cannot have O.N. more than + 5. Therefore O.N. of N in HNO_4 (peroxynitric acid) must be calculated from structure. The structure of HNO_4 is 

It contains one peroxide linkage and therefore, each of the two atoms of peroxide bonds has an O.N. of –1.

Thus, + 1 – 1 – 1 + x – 2 – 2 = 0 or x = + 5.

S40.

$$\overset{+8}{XeO_6^{4-}}(aq) + 2\overset{-1}{F^-}(aq) + 6H^+(aq) \longrightarrow \overset{+6}{Xe}O_3(s) + \overset{0}{F_2}(g) + 3H_2O(l)$$

O.N. of Xe decreases from + 8 (in XeO_6^{4-}) to + 6 (in XeO_3) and that of F increases from –1 (in F^-) to 0 (in F_2). Therefore, XeO_6^{4-} is reduced while F^- is oxidised. The reaction occurs because Na_4XeO_6 (or XeO_6^{4-}) is a stronger oxidising agent than F_2.

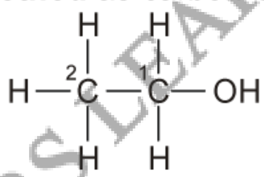
S41. $Ni(CO)_4$: x + 4(0) = 0 or x = 0

$Fe(CO)_5$: x + 5(0) = 0 or x = 5

CH_2O : x + 2(+1) – 2 = 0 or x = 0.

S42. (a) **Conventional Method:** The average oxidation number of carbon in C_2H_5OH or $\overset{x}{C_2}\overset{+1}{H_6}\overset{-2}{O}$ will be

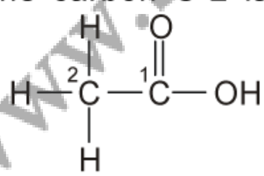
$$2x + 6(+1) + 1(-2) = 0 \quad \text{or} \quad x = -2.$$

On the basis of **chemical bonding method** as carbon C-2 is attached to three H-atoms (less electrongegative than carbon)  and one CH_2OH group (more electronegative than carbon), therefore, O.N. of C-2 will be = 2 (+1) + x + 1 (– 1) = 0 or x = – 2.

The carbon, C-1 is, however, attached to one OH grop (O.N. – 1) and one CH_3 group (O.N. = + 1), therefore, O.N. of C-1 = + 1 + 2 (+ 1) + x + 1 (– 1) = 0 or x = – 2.

(b) **Conventional Method:** The oxidation number of carbon in CH_3COOH or $\overset{x}{C_2}\overset{+1}{H_4}\overset{-2}{O_2}$ will be

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

Chemical bonding method: The carbon C-2 is attached to three H-atoms (less electrongegative 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 group (O.N. – 1) and one CH_3 group (O.N. = + 1), therefore, O.N. of C-1 = + 1 + x + 1 (– 2) + 1 (– 1) = 0 or ***x* = 2**.

S43. (a) In KI_3, since the oxidation number of K is +1, therefore, the avertage oxidation number of iodine = – 1/3. But the oxidation number cannot be fractional. Therefore, in such cases we must consider its structure, $(\overset{+}{K}\overset{-}{I_3})$. In I_3^- ion, 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 forming the coordinate bond is – 1.

(b) **Conventional Method:** The average oxidation number of Fe in $\overset{x}{Fe_3}\overset{-2}{O_4}$ will be

$$3x + 4(-2) = 0 \quad \text{or} \quad x = 8/3.$$

Fe_3O_4 is mixed therefore its oxidation number is calculated by Stochiometry method. Accordingly, Fe_3O_4 shall be written as:

$$\overset{+2}{Fe}\overset{-2}{O} \cdot \overset{-3}{Fe_2}\overset{-2}{O_3}$$

Fe has O.N. of +2 and +3.

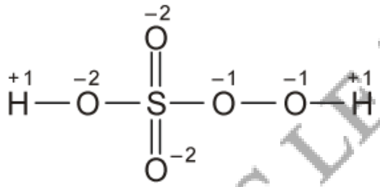
S44. (i) **O.N. of S in H_2SO_5:**

On the basis of conventional method, the O.N. of S in H_2SO_5 comes out to be +8 as shown below.

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

This is impossible because the maximum O.N. of S cannot be more than six since it has only six electrons in the valence shell. This fallacy is overcome by calculating the O.N. of S by *chemical bonding method*.

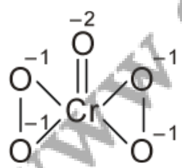
The structure of H_2SO_5 along with O.N. of different atoms is:



Hence, $2(+1) + 3(-2) + x + 2(-1) = 0 \quad \text{or} \quad x = +6.$

(ii) **O.N. of S in CrO_5:**

According to conventional methos, O.N. of Cr is: $x + 5(-2) = 0$ or $x = +10$. This is impossible because the maximum O.N. of Cr cannot be more than six since it has a maximum of six electrons ($3d^5 4s^1$) in the outer orbital configuration which can participate in bonding. This fallacy is removed by calculating O.N. of Cr by *chemical bonding method*. The structure of CrO_5 is:



From the structure, the O.N. of Cr can be calculated as follows:

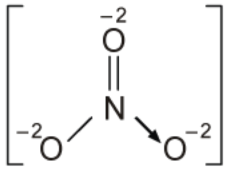
$$x + 4(-1) + 1(-2) = 0 \quad \text{or} \quad x = +6$$

It should be noted that four O atoms have peroxy linkage (O.N. = – 1) and O held by double bond has O.N. = – 2.

(iii) **O.N. of N in NO_3^- :**

According to conventional methos, O.N. of N in $NO_3^- = x + 3(-2) = -1$ or $x + 5$.

According to *chemical bonding method*, the structure of nitrate ion is:



Hence the O.N. of nitrogen is,

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

Thus, the O.N. of N in NO_3^- whether one calculates by conventional method or by chemical bonding method is same.

S45. In the given reaction O.N. of F_2 changes from zero to – 1 in HF and HOF whereas O.N. of oxygen chage from – 2 in H_2O to zero in HOF. Thus, F_2 is reduced, whereas oxygen is oxidised and, therefore, it is a redox reaction.

HOF an highly unstable molecule and hence decomposes to form O_2 and HF

$$2HOF \longrightarrow 2\overset{+1}{H}\overset{-1}{F} + \overset{0}{O_2}$$

If oxygen of HOF disproportionates, then in the disproportionation reaction oxygen must have three oxidation states. In simple words, the oxidation state of oxygen in one of the products should be lower than O in HOF whereas in the other product it should be higher. The oxidation state of oxygen is zero in HOF. It decrease its oxidation state to – 2 if HOF gets reduced to H_2O and also it can increase its oxidation state to + 2 if HOF gets oxidised to OF_2. Therefore, the possible reaction is:

$$2\overset{+1}{H}-\overset{0}{O}-\overset{-1}{F} \longrightarrow \overset{+1}{H_2}\overset{-2}{O} + \overset{+2}{O}\overset{-1}{F_2}$$

S46. Any substance can act as an oxidising or reducing agent if one of the elements in the given compound is present in its intermediate oxidation state. Thus, during the reaction it can either increase its O.N. (reducing agent) or decrease its O.N. (oxidising agent).

(a) In SO_2, ON of *S* is + 4. In principle, *S* can have a minimum O.N. of – 2 and maximum of + 6. Therefore, *S* in SO_2 can either decrease or increase its O.N. and hence can act both as an oxidising as well as a reducing agent.

(b) In H_2O_2, the O.N. of O is – 1. In principle, O can have a minimum O.N. of – 2 and maximum of zero (+ 2 is possible only with OF_2). Therefore, O in H_2O_2 can either decrease its O.N. from – 1 to – 2 or can increase its O.N. from – 1 to zero. Therefore, H_2O_2 acts both as an oxidising as well as a reducing agent.

(c) In O_3, the O.N. of O is zero. It can only decrease its ON from zero to – 1 or – 2, but cannot increaseto + 2. Therefore, O_3 acts only as an oxidant.

(d) In HNO_3, nitrogen atom is in its maximum oxidation state of + 5. Therefore, it can only decrease its O.N. and hence can act as an oxidant only.

S47. The list of substance oxidised, reduced, etc., for the given reactions are tabulated below:

	Substance oxidised	Substance reduced	Oxidising agent	Reducing agent
(a)	$HCHO(l)$	$Cu^{2+}(aq)$	$Cu^{2+}(aq)$	$HCHO(l)$
(b)	$N_2H_4(l)$	$H_2O_2(l)$	$H_2O_2(l)$	$N_2H_4(l)$
(c)	$Pb(s)$	$PbO_2(s)$	$PbO_2(s)$	$Pb(s)$

S48. The Skeletal equation for the process would be:

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

Reduction half equation:

$$Cl_2(aq) \longrightarrow Cl^-(aq)$$

Balance Cl atoms,

$$\overset{0}{Cl_2}(aq) \longrightarrow 2Cl^-(aq)$$

Balance O.N. by adding electrons,

$$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq) \quad \text{... (i)}$$

Oxidation half equation:

$$\overset{+4}{SO_2}(aq) \longrightarrow \overset{+6}{SO_4^{2-}}(aq) + 2e^-$$

Balance O.N. by adding electrons,

$$SO_2(aq) \longrightarrow SO_4^{2-}(aq) + 2e^-$$

Balance O atoms by adding $2H_2O$,

$$SO_2(aq) + 2H_2O(l) \longrightarrow SO_4^{2-}(aq) + 2e^-$$

Balance charge by adding H^+ ions:

$$SO_2(aq) + 2H_2O \longrightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \quad \text{... (iii)}$$

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

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

S49. (a) Fluorine (F) being most electronegative element shows only a –ve oxidation state of –1.

(b) Cesium (Cs) is an highly electropositive alkali metal and has a single electron in the valence shell. Therefore it can exhibit an oxidation state of +1.

(c) Iodine (I) because of the presence of seven electrons in the valence shell, shows an oxidation state of – I and because of the presence of *d*-orbitals it also exhibits +ve oxidation states of +1, +3, +5 and +7.

(d) Neon (Ne) is an inert gas and hence it neigher exhibits –ve nor +ve oxidation state.

S50. In Cu_2O, copper in +1(Cu^+) oxidation state and it undergoes disproportionation to form Cu and Cu^{2+} as :

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

Thus, Cu^+ or Cu_2O acts both as an oxidant as well as a reductant.

For example,

(a) When heated in air, Cu_2O is oxidised to CuO

$$2\overset{+1}{Cu_2}O + \frac{1}{2}O_2 \longrightarrow 2\overset{+2}{Cu}O$$

Thus, Cu_2O acts as a reduactant and reduces O_2 to O^{2-}.

(b) When heated with Cu_2S, it oxidises S^{2-} to SO_2 and hence Cu_2O acts as an oxidant.

$$2\overset{+1}{Cu_2}O + \overset{+1}{Cu_2}S \longrightarrow 6\overset{0}{Cu} + SO_2$$

S51. Amount of $K_2Cr_2O_7$ in 15 mL of 0.02 M solution = 0.02 mol $L^{-1} \times 15 \times 10^{-3}$ L

= 0.03×10^{-3} mol

Let us suppose that the amount of Fe^{2+} ions in 10 mL of solution is *x*.

Redox reaction: $$\underset{1\text{ mol}}{\overset{0.3\times10^{-3}}{Cr_2O_7^{2-}}} + \underset{6\text{ mol}}{\overset{x}{6\,Fe^{2+}}} + 14\,H^+ \rightarrow 2\,Cr^{3+} + 6\,Fe^{3+} + 7H_2O$$

Ratio proportion: $$\frac{x}{6\text{ mol}} = \frac{0.3\times10^{-3}\text{ mol}}{1\text{mol}}$$

$$x = \frac{0.3\times10^{-3}\text{ mol}\times 6\text{ mol}}{1\text{mol}} = 1.8\times10^{-3}\text{ mol}$$

Amount of Fe^{2+} ions in 10×10^{-3} L (10 mL) of solution = 1.8×10^{-3} mol

Amount of Fe^{2+} ions in 1 L of soluiton = $\dfrac{1.8\times10^{-3}\text{mol}}{10\times10^{-3}\text{ L}}$

Molarity = 0.18 mol/L

S52. (a) $2\,Pb(NO_3)_2\,(s) \rightarrow 2\,PbO\,(s) + 4\,NO_2(g) + \frac{1}{2}O_2\,(g)$

It is a decomposition reaction.

Reason: In this reaction a single substance breaks to give three different substances.

(b) $2\,NO_2(g) + 4\,OH^-\,(aq) \rightarrow NO_2^-\,(aq) + NO_3^-\,(aq) + H_2O\,(l)$

It is a disproportionation reaction.

Reason: In this reaction NO_2 (Oxidation number of N = + 4) disproportionates into NO_2^- (oxidation number of N = + 3) and NO_3^- (oxidation number of N = + 5) and

(c) $N_2(g) + O_2\,(g) \rightarrow 2NO\,(g)$

It is a combination reaction.

Reason: In this reaction two different pure substances react to give a single new substance.

Date: 16/10/2021

Q1. Using half reaction method balance the following equations:

$ClO_3^- + Mn^{2+} \longrightarrow MnO_2 + Cl_2$ (in acidic medium)

Q2. Balance the equation for the following reactions by oxidation number method:

$Cr(OH)_4^- + H_2O_2 \longrightarrow Cr_2O_4^{2-} + H_2O$ (basic solution)

Q3. Balance the equation for the following reactions by oxidation number method:

$MnO_4^-(aq) + Fe(OH)(s) \longrightarrow Fe(OH)_2(s) + Cr(OH)_4^-(aq)$

Q4. Balance the equation for the following reactions by oxidation number method:

$MnO_4^-(aq) + CN^-(aq) \longrightarrow CNO^-(aq) + MnO_2(s)$ (in basic solution)

Q5. Balance the equation for the following reactions by oxidation number method:

$MnO_4^{2-}(aq) + SO_3^{2-}(aq) \longrightarrow SO_4^{2-}(aq) + MnO_2(s)$ (in basic solution)

Q6. Balance the equation for the following reactions by oxidation number method:

$Cr_2O_7^-(aq) + Fe^{2+}(aq) + H^+(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq) + H_2O(I)$

Q7. Balance the equation for the following reactions by oxidation number method:

$Fe(s) + O_2(g) + H^+(aq) \longrightarrow Fe^{2+}(aq) + H_2O(I)$

Q8. Balance the equation for the following reactions by oxidation number method:

$IO_3^-(aq) + Cl^-(aq) \longrightarrow Cl_2(g) + I_2(aq)$ (in acidic solution)

Q9. Using half reaction method balance the following equations:

$I^- + O_2 + H_2O \longrightarrow I_2 + OH^-$

Q10. Using half reaction method balance the following equations:

$HNO_2 + MnO_4^- \longrightarrow Mn^{2+} + NO_3^-$ (in acidic solution)

Q11. Using half reaction method balance the following equations:

$CrO_3^- + H_2O_2 \longrightarrow Cr^{6+} + H_2O$ (in basic solution)

Q12. Using half reaction method balance the following equations:

$Cu(s) + Au^+ \longrightarrow Cu^{2+}(aq) + Au(s)$

Q13. Using half reaction method balance the following equations:

$I^- + IO_2^- + H^+ \longrightarrow I_2 + H_2O$

Q14. Using half reaction method balance the following equations:

$MnO_4^- + H_2C_2O_4 \longrightarrow Mn^{2+} + CO_2$ (in acidic solution)

Q15. Using half reaction method balance the following equations:

$Al + NO_3^- \longrightarrow Al(OH)_4^- + NH_3$ (in basic solution)

Q16. Using half reaction method balance the following equations:

$Fe(OH)_2 + H_2O_2 \longrightarrow Fe(OH)_3 + H_2O$ (in basic solution)

Q17. Using half reaction method balance the following equations:

$$NO_2 + I^- \longrightarrow I_2 + NO$$ (in basic solution)

Q18. Using half reaction method balance the following equations:

$$Bi + NO_3^- + H^+ \longrightarrow Bi^{3+} + NO_2 + H_2O$$

Q19. Using half reaction method balance the following equations:

$$H_2O + Fe^{3+} \longrightarrow Fe^{2+} + S + H^+$$ (in basic solution)

Q20. Using half reaction method balance the following equations:

$$PbO_2 + Cl^- \longrightarrow Pb(OH)_3^- + ClO^-$$ (in acidic solution)

Q21. Using half reaction method balance the following equations:

$$Cu + NO_3^- \longrightarrow Cu^{2+} + NO_2$$

Q22. Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidation number methods and identify the oxidising agent and the the reducing agent.

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

Q23. Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the the reducing agent.

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

Q24. Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the the reducing agent.

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

Date: 16/10/2021

S1. $5Mn^{2+} + 2ClO_3^- \longrightarrow 5MnO_2 + 8H^+ + Cl_2$

S2. $2Cr(OH)_4^- + 3H_2O_2 + 2OH^- \longrightarrow 2CrO_4^{2-} + 8H_2O$

S3. $CrO_4^{2-}(aq) + 3Fe(OH)_2(s) \longrightarrow Fe(OH)_3(s) + Cr(OH)_4^-(aq) + OH^-$

S4. $2MnO_4^- + 3CN^- + H_2O \longrightarrow 2MnO_2 + 3CNO^- + 2OH^-$

S5. $MnO_4^{2-}(aq) + SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + MnO_2 + 2OH^-$

S6. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 6\,Fe^{3+}(aq) + 7H_2O$

S7. $2Fe\,(s) + O_2(g) + H^+(aq) \longrightarrow 2Fe^{2+}(aq) + H_2O(I)$

S8. $IO_3^-(aq) + 10Cl^-(aq) + 12H^+(aq) \longrightarrow I_2(ag) + 5\,Cl_2(g) + 6H_2O(I)$

S9. $4I^- + O_2 + 2H_2O \longrightarrow 2I_2 + 4OH^-$

S10. $5HNO_2 + 2MnO_4^- + H^+ \longrightarrow 5NO_3^- + 2Mn^{2+} + 3H_2O$

S11. $2CrO_3^- + 6H_2O + H_2O_2 \longrightarrow 2Cr^{6+} + 14OH^-$

S12. $Cu(s) + 2Au^+(aq) \longrightarrow Au(s) + Cu^{2+}(aq)$

S13. $10I^- + 2IO_3 + 12H^+ \longrightarrow 6I_2 + 6H_2O$

S14. $5H_2C_2O_4 + 2MnO_4^- + 6H^+ \longrightarrow 10CO_2 + Mn^{2+} + 8H_2O$

S15. $8Al + NO_3^- \longrightarrow 5OH^- + 2Fe(OH)_3$

S16. $2Fe(OH)_2 + H_2O_2 \longrightarrow 2Fe(OH)_3$

S17. $2I^- + 3NO_3^- + 5OH^- + 18H_2O \longrightarrow 8Al(OH)_4^- + 3NH_3$

S18. $Bi + 3NO_3^- + 6H^+ \longrightarrow Bi^{3+} + 3NO_2 + 3H_2O$

S19. $H_2S + 2Fe^{3+} \longrightarrow 2Fe^{2+} + S + H^+$

S20. $Cl^- + PbO_2 + 2H^- + H_2O \longrightarrow ClO^- + Pb(OH)_3^-$

S21. $Cu + 2NO_3^- + 4H^+ \longrightarrow 2NO_2 + 2N_2O + Cu^{2+}$

S22.

Oxidation

$$\overset{0}{N_2H_4}(l) + ClO_3^-(aq) \longrightarrow \overset{-3+1}{NO}(g) + \overset{-1}{Cl^-}(g)$$

Reduction

In the reaction, N_2H_4 gets oxidised and acts as a reducing agent and ClO_3^- gets reduced and acts as an oxidising agent.

Oxidation number method

O.N. increases by 4 per N atom or 4×2 per N_2H_4 molecule

$$\overset{-2}{N_2H_4}(l) \quad + \quad ClO_3^- \longrightarrow \overset{+2}{NO}(g) \quad + \quad \overset{+1}{Cl^-}(aq)$$

O.N. decreases by 6 per Cl atom

To balance the increase or decrease in oxidation number, multiply N_2H_4 by 3 and ClO_3^- by 4 and add

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

Balance N and Cl atoms on R.H.S.

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

Balance N and Cl atoms on R.H.S.

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

Balance N and Cl atoms on R.H.S.

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

Since H atoms are balanced automatically, the above equation represents balanced chemical equation.

Ion electron method

Oxidation half reaction: $N_2H_4(l) \longrightarrow NO(g)$

Reduction half reaction: $ClO_3^- \longrightarrow Cl^-(aq)$

These can be balanced separately as:

Oxidation half reaction:

$$N_2H_4(l) \longrightarrow NO(g)$$

$$N_2H_4(l) \longrightarrow 2NO(g)$$

$$\overset{-2}{N_2}H_4(l) \longrightarrow 2\overset{+2}{NO}(g) + 8e^-$$

$$N_2H_4(l) + 8OH^-\ (aq) \longrightarrow 2NO(g) + 8e^- \qquad \text{(balancing OH–ions)}$$

$$N_2H_4(l) + 8OH^-(aq) \longrightarrow 2NO(g) + 6H_2O(l) + 8e^- \qquad \text{... (i)}$$

Reduction half reaction:

$$ClO^-_3 \longrightarrow Cl^-(aq)$$

$$\overset{+5}{ClO_3^-} + 6e^-(s) \longrightarrow \overset{+1}{ClO^-}(aq)$$

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

$$ClO_3^-\ (aq) + 3H_2O(l) + 6e^- \longrightarrow Cl^-(aq) + 6\ OH^-(aq) \quad \ldots \text{(ii)}$$

To balance electrons, multiply eq (i) by 3 and eq (ii) by 4

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

S23.

Reduction

$$\overset{0}{P_4}(s) + OH^-(aq) \longrightarrow \overset{-3+1}{PH_3}(g) + H_2\overset{+1}{P}O_2^-$$

Oxidation

P_4 gets oxidised to $H_2PO_2^-$ and reduced to PH_3 and therefore, P_4 acts as reducing agent as well as oxidising agent.

Oxidation number method

O.N. decreases by 3 per atom or 3×4 per P_4 molecule

$$\overset{0}{P_4}(s) + OH^-(aq) \longrightarrow 3\overset{-3+1}{PH_3}(g) + H_2\overset{+1}{P}O_2^-(aq)$$

O.N. increases by 1 per atom or 1×4 per P_4 molecule

Balance the increase or decrease in oxidation number by multiplying $H_2PO_2^-$ by 3 and PH_3 by 1.

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

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

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

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

$P_4(s) + 6OH^-(aq) + 3H_2O\ (l) \longrightarrow PH_3(g) + 3H_2PO_2^-\ (aq)$ To balance H atoms, add 3 H_2O to L.H.S and $3OH^-$ on R.H.S, we have

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

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

Ion electron method

The equation can be split up as

Oxidation half reaction: $P_4(s) \longrightarrow PH_3(g)$

These can be balanced separately as :

Oxidation half reaction

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

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

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

$$P_4(s) + 8OH^-(aq) \longrightarrow 4H_2PO_2^-(aq) + 4e^- \qquad \ldots \text{(i)}$$

Reducing half reaction

$$\overset{0}{P_4} \longrightarrow \overset{-3}{P}H_3(g)$$

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

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

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

$$P_4(s) + 12H_2O(l) + 12e^- \longrightarrow 4PH_3(g) + 12OH^-(aq) \qquad \ldots \text{(ii)}$$

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

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

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

S24.

Reduction

$$\overset{+7}{Cl_2O_7}(g) + \overset{-1}{H_2O_2}(aq) \longrightarrow \overset{+3}{ClO_2^-}(aq) + \overset{0}{O_2}(g) + H^+$$

Oxidation

In this reaction, Cl_2O_7 gets reduced acts as an oxidising agent while H_2O_2 gets oxidise and acts as a reducing agent.

Oxidation number method

O.N. decreases by 4 per atom or 4 × 2 per molecule

$$\overset{+7}{Cl_2O_7}(g) + \overset{-1}{H_2O_2}(aq) \longrightarrow \overset{+3}{ClO_2^-}\ (aq) + \overset{0}{O_2}(g)$$

O.N. increases by 1 per atom or 1 × 2 per molecule

To balance the increase decrease in oxidation number multiply H_2O_2 by 4 and add

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

Balance Cl atoms and O_2

$$Cl_2O_7(g) + 4H_2O(aq) \longrightarrow ClO_2^-\ (aq) + 4O_2(g)$$

To balance Cl multiply ClO_2^- by 2.

$$Cl_2O_2(g) \ + \ 4H_2O_2(aq) \longrightarrow 2ClO_2^-(aq) \ + \ 4O_2(g)$$

To balance O atoms, add $3H_2O$ molecules on R.H.S.

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

To balance O atoms, add $3H_2O$ molecules on R.H.S.

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

To balance H atoms, add $2H^+$ on R.H.S. since the medium is acidic

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

Ion electron method

Oxidation half reaction : $H_2\overset{-1}{O_2}(aq) \longrightarrow \overset{0}{O_2}(g)$

Reduction half reaction : $\overset{+7}{Cl_2}O_7(g) \longrightarrow \overset{+5}{Cl}O_3^-(aq)$

These can be balanced separately as :

Oxidation half reaction

$$H_2\overset{-1}{O_2}(ag) \longrightarrow O_2(g) + 2e^-$$

Since the reaction occurs in acidic medium add $2H^+$ on R.H.S. to equate charge,

$$H_2O_2(aq) \longrightarrow O_2(g) + 2e^- + 2H^+(aq) \quad \text{... (i)}$$

Reduction half reaction

$$Cl_2O_7(g) \longrightarrow ClO_2^-(aq)$$

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

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

$$Cl_2O_7(g) + 8e^- + 6H^+(aq) \longrightarrow 2ClO_2^-(aq) \quad \text{(medium is acidic)}$$

$$Cl_2O_7(g) + 6H^+(aq) + 8e^- \longrightarrow 2ClO_2^-(aq) + 3H_2O \quad \text{... (ii)}$$

To balance electrons multiply eq.(i) by 4 and add to eq (ii)

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