Topic : **<u>Redox reactions</u>** **F.Y.B.Sc**

Q.1	**Define the following terms:** **a)Oxidation b) Reduction c) Oxidation number d) Oxidising agent e) Reducing agent f) Redox reaction**	
Ans.	a)**Oxidation** : It is defined as a process in which an atom or an ion losses one or more electrons. b)**Reduction** : It is defined as a process in which an atom or an ion accepts one or more electrons. c)**Oxidation number or oxidation state**: It is defined as an apparent charge present on an atom or an ion of an element in a compound **OR** Oxidation number or oxidation state of an atom in a molecule or ion is defined as the number of charges it would carry if electrons were completely transferred. d) **Oxidising agent** : It is defined as a species that accept electron(s) and causes other substance to lose electron(s). **OR** Oxidising agent is defined as a substance which oxidizes other substance and itself get reduced by gain of electrons. e) **Reducing agent :** It is defined as a species that losses electron(s) and causes other substance to accept electron(s) **OR** Reducing agent may also be defined as a substance which reduces other substance and itself gets oxidized by loss of electro(s). f)**Redox reaction** : The chemical reactions in which oxidation and reduction reactions occur simultaneously are called redox reaction.	
Q.2	**Distinguish between oxidation reaction and reduction reaction.**	
	Oxidation	**Reduction**
	By old concept:	
	i) Addition of oxygen ii) Removal of hydrogen iii) Addition of electronegative element iv) Removal of electropositive element	i) Addition of hydrogen ii) Removal of oxygen iii) Addition of electropositive element iv) Removal of electronegative element
	By modern electronic concept:	
	i)Lose of one or more electrons ii)Increase in oxidation number or positive charge	i)gain of one or more electrons ii)Decrease in oxidation number or positive charge

Q.3	**Give the characteristics of oxidising agents and reducing agent.**
Ans.	**Characteristics of oxidising agent** (oxidant) a)It accepts electrons b)It itself undergo reduction c)It causes oxidation d)It undergoes decrease in oxidation number **Characteristics of reducing agent** (reductant) a)It loses electrons b)It itself undergo oxidation c)It causes reduction d)It undergoes an increase in oxidation number
Q.4	**Give the rules to assign the oxidation number of an element.**
Ans.	**a)**The oxidation number of an element in free or uncombined state is zero **b)**The oxidation number of alkali metals (ie.Li, Na, K, Rb, Cs , Fr is alwys +1) **c)**The oxidation number of alkaline earth metals ie.Mg, Ca, Sr, Ba and Ra is always +2 **d)**The oxidation number of an atom in a monoatomic ion is equal to its charge. **e)**The oxidation number of oxygen in all oxides is ‘ – 2’ but in peroxides like BaO_2 , H_2 , O_2 etc. and in superoxides like KO_2 etc. it is ‘ – 1’ **f)**The oxidation number of hydrogen is +1 but in metallic hydrides like CaH, NaH, BeH_2 etc. it is ‘ – 1 ’. **g)**The oxidation number of fluorine is always ‘- 1’ in all its compounds, while other halogens like Cl, Br , I along with ‘ – 1’ oxidation number, they may have positive oxidation state. **h)**The algebraic sum of all oxidation states of all atoms in a neutral molecule is zero while for a polyatomic ion, it is equal to the net charge on the ion.
Q.5	**Identify the oxidising agents and reducing agents from the following equations:** **i)**$H_2C_2O_4 + MnO^-_{4(aq.)} \rightarrow CO_{2(g)} + Mn^{2+}$ **ii)**$Sn^{2+}_{(aq.)} + IO^-_{4(aq.)} \rightarrow Sn^{4+}_{(aq.)} + I^-_{(aq.)}$ **iii)**$H_2O_{2(aq.)} + Cr_2O_7^{-2}{}_{(aq.)} \rightarrow Cr^{3+}_{(aq.)} + O_{2(g)}$ **iv)**$FeCl_{2(aq.)} + 2KOH_{(aq.)} \rightarrow Fe(OH)_{2(aq.)} + 2KCl_{(aq.)}$
Ans.	i) $H_2C_2O_4 + MnO^-_{4(aq.)} \rightarrow CO_{2(g)} + Mn^{2+}$ +1 +3 -2 +7 -2 +4 -2 +2 Gain of electrons loss of e^- Increase in oxidation number of carbon occurs by loss of 2 electrons. Hence $H_2C_2O_4$ behaves as reducing agent. Decrease in oxidation number of Mn occurs by gain of 5 electrons. Hence it acts as

	oxidising agent. ii) $Sn^{2+}{}_{(aq.)} + I\ O^{-}{}_{4(aq.)} \rightarrow Sn^{4+}{}_{(aq.)} + I^{-}{}_{(aq.)}$ ↓ ↓ ↓ ↓ ↓ +2 +7 - 2 +4 - 1 gain of electrons Loss of electrons Increase in oxidation number of Sn^{2+} to Sn^{4+} occurs by loss of 2 electrons. Hence Sn^{2+} is reducing agent. Decrease in oxidation number of Iodine occurs from +7 to - 1 by gain of 8 electrons. Hence IO_4^{-} acts as oxidising agent. iii) $H_2O_{2\ (aq.)} + Cr_2O_7^{-2}{}_{(aq.)} \rightarrow Cr^{3+}{}_{(aq.)} + O_{2(g)}$ ↓ ↓ ↓ ↓ ↓ ↓ +1 - 1 +6 -2 +3 0 Gain of electrons Loss of electrons Increase in oxidation number of oxygen from -1 to 0 occurs by loss of one electron , hence H_2O_2 is a reducing agent. Decrease in oxidation number of Cr from - 1 to 0 occurs by gain of 6 electrons. Hence $Cr_2O_7^{-2}$ acts as oxidising agent. iv) $Fe\ Cl_{2(aq.)} + 2\ K\ O\ H_{(aq.)} \rightarrow Fe\ (O\ H)_{2(aq.)} + 2K\ Cl_{(aq.)}$ ↓ ↓ ↓ ↓ ↓ ↓ ↓ ↓ ↓ ↓ +2 - 1 +1 -2 +1 +2 -2 +1 +1 -1 As there is no change in oxidation number, the above reaction is not a redox reaction. Hence there is no oxidising agent and reducing agent.
Q.6	**Find the oxidation states of the underlined elements:** **a)$\underline{C}H_4$ b) $\underline{Cr_2}O_7^{2-}$ c) $\underline{Ba}O_2$ d) $\underline{N}H_4^{+}$ e) $K_3[\underline{Fe}(OH)_6]$**
Ans.	a) Let the oxidation number of carbon be x x+4(1) = 0 therefore, x = -4 Hence oxidation state of carbon is ‘-4’ b) Let the oxidation number of Cr be x 2x + 7(-2) = -2 , 2x – 14 = -2 2x = 12 x = 6 Therefore, oxidation state of chromium is +6 c) Let the oxidation state of Ba be x As it is a peroxide, oxidation sate of oxygen is ‘- 1’

	Therefore, $x + 2(-1) = 0$ $X = 2$ Therefore, oxidation state of Ba is +2 d) Let the oxidation number of nitrogen be x $x + 4(1) = +1$ $x + 4 = 1, \quad x = -3$ Therefore, oxidation state of nitrogen '- 3' e) Let the oxidation state of Fe be x oxidation state of CN^- is '- 1' therefore, $3(1) + x + 6(-1) = 0$ $3 + x - 6 = 0, \quad x = 3$ Therefore, oxidation state of Fe is +3
Q.8	**Balance the following equations by ion-electron method.** **i)** $SO_{2(g)} + Fe^{3+}_{(aq.)} \rightarrow Fe_{(aq.)}^{2+} + SO_{4(aq.)}^{2-}$ **(acidic)** **ii)** $ClO^-_{(aq.)} + Cr(OH)_{3(aq.)} \rightarrow CrO_{4(aq.)}^{2-} + Cl^-_{(aq.)}$ **(basic)** **iii)** $H_2C_2O_{4(aq.)} + MnO_4^-{}_{(aq.)} \rightarrow CO_{2(g)} + Mn^{2+}_{(aq.)}$
Ans.	**Step I** Assign oxidation number to each element $\underset{+4\ -2}{SO_{2(g)}} + \underset{+3}{Fe^{3+}_{(aq.)}} \rightarrow \underset{+2}{Fe_{(aq.)}^{2+}} + \underset{+6\ -2}{SO_{4(aq.)}^{2-}}$ **Step II** Divide the equation into two half equations, an oxidation half and reduction half $SO_{2(g)} \rightarrow SO_{4(aq.)}^{2-}$ oxidation $Fe^{3+}_{(aq.)} \rightarrow Fe_{(aq.)}^{2+}$ reduction **Step III** Balance the half equation for O atoms adding H_2O to the side with less O atoms $SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_{4(aq.)}^{2-}$ oxidation $Fe^{3+}_{(aq.)} \rightarrow Fe_{(aq.)}^{2+}$ reduction **Step IV** Balance the H atoms by adding H^+ ions to the side having less H atoms $SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_{4(aq.)}^{2-} + 4H^+$ oxidation $Fe^{3+}_{(aq.)} \rightarrow Fe_{(aq.)}^{2+}$ reduction **Step V** Add appropriate number of electrons for oxidation reaction, add electrons on R.H.S. while for reduction, add electrons on L.H.S. $SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_{4(aq.)}^{2-} + 4H^+ + 2e^-$ oxidation $Fe^{3+}_{(aq.)} + e^- \rightarrow Fe_{(aq.)}^{2+}$ reduction **Step VI** Balance the electron by multiplying with a suitable factor. $SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_{4(aq.)}^{2-} + 4H^+ + 2e^- + 6e^-$ oxidation

$2Fe^{3+}_{(aq.)} + 2e^- \rightarrow 2Fe^{2+}_{(aq.)}$ reduction

Step VII Add both the equations

$SO_{2(g)} + 2H_2O_{(l)} + 2Fe^{3+}_{(aq.)} \rightarrow SO_{4(aq.)}^{2-} + 4H^+ + 2Fe^{2+}_{(aq.)}$

ii)

Step I) Assign oxidation number to each element

$ClO^-_{(aq.)} + Cr(OH)_{3(aq.)} \rightarrow CrO_{4(aq.)}^{2-} + Cl^-_{(aq.)}$

↓ ↓ ↓ ↓ ↓ ↓ ↓ ↓

+1 -2 +3 -2 +1 +6 -2 -1

Step II Divide the equation into two half equations, an oxidation half and reduction half

$Cr(OH)_{3(aq.)} \rightarrow CrO_{4(aq.)}^{2-}$ oxidation

$ClO^-_{(aq.)} \rightarrow ClO^-_{(aq.)}$reduction

Step III Balance the half equation for O atoms adding H_2O to the side with less O atoms

$Cr(OH)_{3(aq.)} + H_2O_{(l)} \rightarrow CrO_{4(aq.)}^{2-}$ oxidation

$ClO^-_{(aq.)} \rightarrow Cl^-_{(aq.)} + H_2O$ reduction

Step IV Balance the H atoms by adding H^+ ions to the side having less H atoms

$Cr(OH)_{3(aq.)} + H_2O_{(l)} \rightarrow CrO_{4(aq.)}^{2-} + 5H^+$ oxidation

$ClO^-_{(aq.)} + 2H^+ \rightarrow Cl^-_{(aq.)} + H_2O$ reduction

Step V Add appropriate number of electrons for oxidation reaction, add electrons on R.H.S. while for reduction, add electrons on L.H.S.

$Cr(OH)_{3(aq.)} + H_2O_{(l)} \rightarrow CrO_{4(aq.)}^{2-} + 5H^+ + 3e^-$ oxidation

$ClO^-_{(aq.)} + 2H^+ + 2e^- \rightarrow Cl^-_{(aq.)} + H_2O$ reduction

Step VI Balance the electron by multiplying with a suitable factor.

$2Cr(OH)_{3(aq.)} + 2H_2O_{(l)} \rightarrow 2CrO_{4(aq.)}^{2-} + 10H^+ + 3e^-$ oxidation

$3ClO^-_{(aq.)} + 6H^+ + 6e^- \rightarrow 3Cl^-_{(aq.)} + 3H_2O$ reduction

Step VII Add both the equations

$2Cr(OH)_{3(aq.)} + 3ClO^-_{(aq.)} + 6H^+ \rightarrow 2CrO_{4(aq.)}^{2-} + 3Cl^-_{(aq.)} + 6H_2O$

Step VIII As the medium is basic add OH^- ions equal to the number of H^+ on both sides

$2Cr(OH)_{3(aq.)} + 3ClO^-_{(aq.)} + 6H^+ + 6OH^-_{(aq.)} \rightarrow 2CrO_{4(aq.)}^{2-} + 3Cl^-_{(aq.)} + 6H_2O + 6OH^-$

iii)

Step I Write the oxidation number of each element

$$H_2C_2O_{4(aq.)} + MnO_4^{-}{}_{(aq.)} \rightarrow CO_{2(g)} + Mn^{2+}{}_{(aq.)}$$

↓ ↓ ↓ ↓ ↓ ↓ ↓ ↓

+1 +3 -2 +7 -2 -4 -2 +2

Step II

Divide the equation into two half equations, an oxidation half and reduction half and balance the carbon atom in the initial stage

$H_2C_2O_{4(aq.)} \rightarrow 2CO_{2(g)}$ oxidation

$MnO_4^{-}{}_{(aq.)} \rightarrow Mn^{2+}{}_{(aq.)}$ reduction

Step III Balance the half equation for O atoms adding H_2O to the side with less O atoms

$H_2C_2O_{4(aq.)} \rightarrow 2CO_{2(g)}$ oxidation

$MnO_4^{-}{}_{(aq.)} \rightarrow Mn^{2+}{}_{(aq.)} + 4H_2O_{(l)}$ reduction

Step IV Balance the H atoms by adding H^+ ions to the side having less H atoms

$H_2C_2O_{4(aq.)} \rightarrow 2CO_{2(g)} + 2H^+$ oxidation

$MnO_4^{-}{}_{(aq.)} + 8H^+ \rightarrow Mn^{2+}{}_{(aq.)} + 4H_2O_{(l)}$ reduction

Step V Add appropriate number of electrons for oxidation reaction, add electrons on R.H.S. while for reduction, add electrons on L.H.S.

$MnO_4^{-}{}_{(aq.)} + 8H^+ + 5e^- \rightarrow Mn^{2+}{}_{(aq.)} + 4H_2O_{(l)}$ reduction

Step VI Balance the number of electrons by multiplying the equations with appropriate suitable factor.

$5H_2C_2O_{4(aq.)} \rightarrow 10CO_{2(g)} + 10H^+ + 10e^-$ oxidation

$2MnO_4^{-}{}_{(aq.)} + 16H^+ + 10e^- \rightarrow 2Mn^{2+}{}_{(aq.)} + 8H_2O_{(l)}$ reduction

Step VII Add both the equations

$5H_2C_2O_{4(aq.)} + 2MnO_4^{-}{}_{(aq.)} + 16H^+ \rightarrow 10CO_{2(g)} + 10H^+ + 2Mn^{2+}{}_{(aq.)} + 8H_2O_{(l)}$

ie. $5H_2C_2O_{4(aq.)} + 2MnO_4^{-}{}_{(aq.)} + 6H^+ \rightarrow 10CO_{2(g)} + 2Mn^{2+}{}_{(aq.)} + 8H_2O_{(l)}$