Long Answer Type Questions

Q. 1. Using stock notation, represent the following compounds: $HAuCl_4$, TI_2O, FeO, Fe_2O_3, CuI, CuO, MnO and MnO_2.

Compound	**Oxidation Number of Metal**	**Stock Notation**
$HAuCI_4$	$\mathbf{1 + \times + 4(-1) = 0}$ $\mathbf{1 + \times - 4 = 0}$ $\mathbf{\times - 3 = 0}$ $\mathbf{\times = +3}$	**HAu (III) CI_4**

Ans. First of all the oxidation number of each metallic element in its compound is calculated. Then these compounds are represented as stock notation:

Compound	**Oxidation Number of Metal**	**Stock Notation**
TI_2O	$2(\times) + (-2) = 0$ $2\times - 2 = 0$ $2\times = 2$ $\times = \frac{2}{2} = +1$	TI_2 (I) O_2
FeO	$\times + (-2) = 0$ $\times - 2 = 0$ $\times = +2$	Fe(II) O
Fe_2O_3	$2(\times) + 3(-2) = 0$ $2\times - 6 = 0$ $2\times = 6$ $\times = \frac{6}{2} = +3$	Fe_2 (III) O_3
CuI	$\times + (-1) = 0$ $\times - 1 = 0$ $\times = +1$	Cu (I) I
CuO	$\times + (-2) = 0$ $\times - 2 = 0$ $\times = +2$	Cu (II) O
MnO	$\times + (-2) = 0$ $\times - 2 = 0$ $\times = +2$	Mn (II) O

MnO_2	$x + 2(-2) = 0$ $x - 4 = 0$ $x = +4$	$Mn(IV)\ O_2$

Q. 2. Calculate the oxidation number of the underlined element in each species:

(i) $\underline{V}O_2{}^{+}$ **(ii)** $\underline{U}O_2{}^{2+}$ **(iii)** $Ba_2\underline{Xe}O_6$

(iv) $K_4\underline{P_2}O_7$ **(v)** K_2S

Ans. (i) $\underline{V}O_2{}^{+}$

Let, oxidation number of V = x

$\therefore$ $x + 2(-2) = 1$

$x - 4 = 1$

$x = 1 + 4 = +5$

(ii) $\underline{U}O_2{}^{2+}$

Let, oxidation number of U = x

$\therefore$ $x + 2(-2) = 2$

$x - 4 = 2$

$x = 2 + 4 = +6$

(iii) $Ba_2\underline{Xe}O_6$

Let, oxidation number Xe = x

$\therefore$ $2(2) + x + 6(-2) = 0$

$4 + x - 12 = 0$

$x - 8 = 0$

$x = +8$

(iv) $K_4\underline{P_2}O_7$

Let, oxidation number of P = x

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

$4 + 2x - 14 = 0$

$2x - 10 = 0$

$2x = 0$

$X = \frac{10}{2} = +5$

(v) K_2S

Let, oxidation number of S = x

∴ $2(1) + x = 0$

$x = -2$

Q. 3. Calculate the oxidation number of the underlined elements in the following compounds:

(i) $\boldsymbol{K_2\underline{Cr}O_4}$ **(ii)** $\boldsymbol{K_2\underline{Cr_2}O_7}$ **(iii)** $\boldsymbol{\underline{Cr_2}O_2CI_2}$

(iv) $\boldsymbol{\underline{Cr_2}(SO_4)_3}$ **(v)** $\boldsymbol{\underline{M}nO\bar{4}}$

Ans. (i) Let x be the oxidation no. of chromium in K_2CrO_4

O.N. of K = +1

O.N. of O = – 2

∴ $2 \times (+1) + x + 4 \times (-2) = 6$

or $x - 6 = 0$ or $x = +6$

Hence, oxidation no. of Cr in K_2CrO_4 = 6.

(ii) Let x be the oxidation no. of Cr in $K_2Cr_2O_7$.

Oxidation no. of K = +1

Oxidation no. of O = – 2

∴ $2 \times (+1) + 2x + 7 \times (-2) = 0$

or $2x - 12 = 0$

or $x = +6$

Hence, oxidation no. of Cr in $K_2Cr_2O_7$

= +6.

(iii) Let x be the oxidation no. of Cr in CrO_2CI_2.

Oxidation no. of CI = – 1

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

or $x - 6 = 0$

or $x = +6$

Hence oxidation no. Cr in CrO_2CI_2

= + 6.

(iv) Let x be the oxidation no. of Cr in $Cr_2(SO_4)_3$.

Oxidation no. of SO_4^{2-} = – 2

∴ $2x + 3 \times (-2) = 0$

or $2x = +6$

or $x = +3$

Hence oxidation no. of Cr in $Cr_2(SO_4)_3$

= + 3.

(v) Let, x be the oxidation no. of Mn in $MnO\bar{4}$

Oxidation of O = – 2

$\therefore \quad x + 4 \times (-2) = -1$

$x - 8 = -1$

$x = -1 + 8 = +7$

Q. 4. (i) Which of the following species, do not show disproportionation reaction and why?

ClO^-, ClO_2^-, ClO_3^- and ClO_4^-

Also write reaction for each of the species that disproportionate.

(ii) Give one example each of the following redox reaction:
(a) Combination reaction
(b) Decomposition reaction
(c) Metal displacement reaction

Ans. (i) Among the oxoanions of chlorine, ClO_4^- does not disproportionate because in this oxoanion. Chlorine is present in its highest oxidation state i.e. +7.
The disproportionation reactions for other there oxoanions of chlorine are as follows:

$$3\overset{+1}{CIO^-} \rightarrow 2\overset{-1}{CI} + \overset{+5}{CIO^-}_3$$

$$6\overset{+3}{CIO^-}_2 \overset{hr}{\rightarrow} 4\,\overset{+5}{CIO^-}_3 + 2\,\overset{-1}{CI} -$$

$$4\overset{+5}{CIO^-}_3 \rightarrow \overset{-1}{CI} - + 3\,\overset{+7}{CIO^-}_4$$

(ii) (a) Example of combination reaction

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

(b) Example of Decomposition reaction:

$2NaH(s) \overset{\Delta}{\rightarrow} 2Na(s) + H_2(g)$

(c) Example of metal displacement reaction:

$V_2O_5(s) + 5Ca(s) \overset{\Delta}{\rightarrow} 2V(s) + 5CaO(s)$

Q. 5. Consider the cell reaction of electrochemical cell : $Ni(s) + 2Ag^+ (aq) \rightarrow Ni^{2+} (aq) + 2Ag(s)$ and answer the following questions:
(i) Write anode and cathode half reactions.
(ii) Mention the direction of flow of electrons.
(iii) How is the electrical neutrality maintained in the solutions of the two half cells?
(iv) Write the formula for calculating standard emf of this cell.
(v) How does the emf change when the concentration of silver ions decreases?
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Ans. For reaction:

$Ni(s) + 2Ag^{+}(aq) \rightarrow Ni^{2+}(aq) + 2Ag(s)$

(i) Oxidation half reaction (at anode)

$Ni(s) \rightarrow Ni^{2+} + (aq) + 2e^{-}$

Reduction half reaction (at cathode)

$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

(ii) The flow of electrons takes place from anode to cathode.

(iii) Salt bridge is used to maintain electrical neutrality in the solution of the two half cells.

(iv) $E°_{cell} = E°_{cathode} - E°_{anode}$

(v) When the concentration of the cell also decreases.

Q. 6. (i) Suggest a scheme of classification of the following redox reactions:

(a) $N_2(g) + O_2(g) \rightarrow 2NO(g)$
(b) $2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 2NO_2(g) + 1/2O_2(g)$
(c) $NaH(s) + H_2O(I) \rightarrow NaOH(aq) + H_2(g)$
(d) $2NO_2(g) + 2OH^{-}(aq) \rightarrow NO_2^{-}(aq) + NO_3^{-}(aq) + H_2O(I)$

(ii) Why do the following reactions proceed differently?
$Pb_3O_4 + 8HCI \rightarrow 3PbCI_2 + CI_2 + 4H_2O$

And

$Pb_3O_4 + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + PbO_2 + 2H_2O$

Ans. (i) (a) In this reaction, nitric oxide is formed by combination of the elemental substances, nitrogen and oxygen. So, it is an example of **combination redox reaction.**

(b) It involves breaking down of lead nitrate into three components. So, it is an example of **decomposition redox reaction.**

(c) In this reaction, hydrogen of water has been displaced by hydride ion into dihydrogen gas. So, it is an example of displacement redox reaction.

(d) This reaction involves disproportionation of NO_2 (+4 state) in to NO_2^{-} (+3 state) and NO_3^{-} (+5 state). So, it is an example of **disproportionation redox reaction.**

(ii) Pb_3O_4 is actually a stoichiometric mixture of 2 mol of PbO and 1 mole of PbO_2. In PbO_2, lead is present in +4 oxidation state, whereas the stable oxidation state of lead in PbO is +2. PbO_2 thus can act as an oxidant and so, it can oxidise Cl^{-} ion of HCl into chlorine. As we know that PbO is a basic oxide. So, the reaction.

$Pb_3O_4 + 8HCI \rightarrow 3PbCI_2 + CI_2 + 4H_2O$ can be splitted into two reactions-
$2PbO + 4HCI \rightarrow 2PbCI_2 + 2H_2O$ (Acid-base reaction)

$$\overset{+4}{PbO_2} + 4\overset{-1}{HCI} \rightarrow \overset{+2}{PbCI_2} + \overset{o}{CI_2} + 2H_2O$$ (Redox Reaction)

Since, HNO_3 itself is an oxidizing agent, so, it is unlikely that the reaction may occur between PbO_2 and HNO_3.However, the acid-base reaction occurs between PbO and HNO_3 as :

$2PbO + 4HNO_3 \rightarrow 2Pb(NO_3)_2 + 2H_2O + 2H_2O$
It is the passive nature of PbO_2 against HNO_3, that makes the reaction different from the one that follows with HCl.