Level - 01

Answer. 1 The total of positive and negative charge should be zero in the compound.

In $X_3(YZ_4)_2$, we have, $3 \times (+2) + 2\{+5 + 4(-2)\}$

$\Rightarrow$ $+6 + 2(5-8) = +6 + 2(-3) = 6 - 6 = 0$

Answer. 2 The element with atomic number 13 will show the oxidation state of (+3) by loosing 3 electrons to attain the stable electronic configuration of nearest noble gas Neon.

Answer. 3 Let the oxidation state of P in $Mg_2P_2O_7$ be x.

Then, $2(+2) + 2x + 7(-2) = 0$

$\Rightarrow$ $4 + 2x - 14 = 0$

$\Rightarrow$ $2x - 10 = 0 \Rightarrow \boxed{x = +5}$

Answer. 4 To show oxidising behavior, H_2SO_4 should get reduced

In reaction, $2HI + H_2SO_4 \longrightarrow I_2 + SO_2 + 2H_2O$

Reduction half: $H_2SO_4 \longrightarrow SO_2$

(S in +6 oxidation state) (S in +4 oxidation state.)

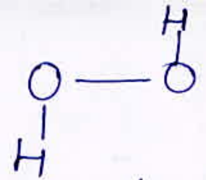
So, H_2SO_4 is getting reduced in this reaction.

Answer. 5 Let oxidation number of S in Na_2SO_4 be x.

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

$\Rightarrow$ $2 + x - 8 = 0 \Rightarrow x - 6 = 0$

$\Rightarrow$ $\boxed{x = +6}$

Answer. 6 The structure of H_2O_2 is . Considering both the O-H bonds to be 100% ionic, we get the oxidation state of both the oxygens as $\underline{-1}$.

Answer. 7 In $C + H_2O \longrightarrow CO + H_2$,

<u>Reduction half:</u> $H_2O \longrightarrow H_2$

(O.N. of $H = +1$) (O.N. of $H = 0$)

~~So,~~ Since H_2O is getting reduced, hence it is working as an <u>oxidising agent.</u>

Answer. 8 $M^{3+} \longrightarrow M^{6+} + 3e^-$

Answer. 9 Let oxidation number of Fe be x.

Then, $3(+1) + \{x + 6(-1)\} = 0$

$\Rightarrow 3 + x - 6 = 0 \Rightarrow x - 3 = 0$

$\Rightarrow \boxed{x = +3}$

Answer. 10 Let oxidation number of S be x. Then,

$2x - 2 = 0 \Rightarrow \boxed{x = +1}$

Answer. 11 Let O.N. of S be x. Then,

$2x - 2(-2) = -2$

$\Rightarrow 2x - 4 = -2 \Rightarrow 2x = -2 + 4$

$\Rightarrow 2x = 2 \Rightarrow \boxed{x = +1}$

Answer. 12 Let O.N. of N be x. Then, $x + 3(+1) = 0$

$\Rightarrow \boxed{x = -3}$

$\boxed{\text{Answer. 13}}$ In acidic medium, the reaction of $K_2Cr_2O_7^{2-}$ is,

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

So, Equivalent weight of $K_2Cr_2O_7 = \frac{MW}{6}$

$\boxed{\text{Answer. 14}}$ In reaction, $\underset{(O.N_p = 0)}{4P} + 3KOH + 3H_2O \longrightarrow \underset{(O.N. = +1)}{3KH_2PO_2} + \underset{(O.N. = -3)}{PH_3}$

So, Phosphorus is both oxidized and reduced.

$\boxed{\text{Answer. 15}}$ Let O.N. of N be x. Then,

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

$$\Rightarrow \boxed{x = -1}$$

$\boxed{\text{Answer. 16}}$ Let oxidation number of P be x. Then,

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

$$\Rightarrow 3 + x - 4 = 0 \Rightarrow \boxed{x = +1}$$

$\boxed{\text{Answer. 17}}$ $\underset{(O.N. = 0)}{3Cl_2} + 6OH^- \longrightarrow \underset{(O.N = -1)}{5Cl^-} + \underset{(O.N = +5)}{ClO_3^-} + 3H_2O$

So, chlorine is oxidised as well as reduced.

$\boxed{\text{Answer. 18}}$ $\underset{(O.N = 0)}{3Br_2} + 6CO_3^{2-} + 3H_2O \longrightarrow \underset{(O.N = -1)}{5Br^-} + \underset{(O.N = +5)}{BrO_3^-} + 6HCO_3^-$

So, bromine is both reduced and oxidised.

$\boxed{\text{Answer. 19}}$ $4Fe + 3O_2 \longrightarrow 4Fe^{3+} + 6O^{2-}$

Metallic iron oxidises to Fe^{3+} and not reduced.

So, option (2) is wrong.

Answer. 20 Oxidation number of Ni in $Ni(CO)_4$ is $\underline{0}$.
because CO is not a radical but a neutral molecule

~~Answer. 21 Let the oxidation numb~~

Answer. 21 Evaluation should be made separately for NH_4^+ and NO_3^-.

Let O.N. of Nitrogen in NH_4^+ and NO_3^- be x and y respectively.

In NH_4^+, $x + 4(+1) = -1 \Rightarrow \boxed{x = -3}$

In NO_3^-, $x + 3 \times (-2) = -1 \Rightarrow \boxed{x = +5}$

Answer. 22 In acidic medium, $Mn^{+7} + 5e^- \longrightarrow Mn^{+2}$
So, it is reduction by five electrons.

Answer. 23 F always shows oxidation state of -1.

Answer. 24 $MnO_4^- + NO_2^- \longrightarrow NO_3^- + Mn^{+2}$

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

Oxidation half: $NO_2^- \longrightarrow NO_3^- + 2e^-$

One mole of MnO_4^- requires 5 mole of electrons and one mole of NO_2^- gives 2 mole of electrons.
So, one mole of MnO_4^- oxidises $\underline{\left(\frac{5}{2}\right)}$ moles of NO_2^-.

Answer. 25 $As_2S_5 + NO_3^- \longrightarrow AsO_4^{3-} + SO_4^{2-} + NO_2$

Oxidation number 'S' ~~to~~ changes from -2 (in As_2S_5) to $+6$ (in SO_4^{2-}). So, one S atom exchange 8 electrons, and 5 S atoms exchange (5 × 8) electrons. n-factor of $As_2S_5 = 40$

So, equivalent weight of $As_2S_5 = \frac{M}{40}$

Answer. 26 Equivalent weight $= \frac{M.W}{n\text{-factor}} = \frac{M.W}{3} \Rightarrow \boxed{n\text{-factor} = 3}$

So, the reaction must be $K\underset{n\text{-factor}=3}{\underline{Mn}}O_4 \longrightarrow Mn^{+4}$

So, the oxidation state of Mn in the final product is $+4$.

Answer. 27 Since **oxidation** number of F is always -1, therefore in ClF_3 oxidation number of Cl must be $+3$.

Answer. 28 Let the oxidation number of Co be x. Then,

$$x + 6(-1) = -3 \Rightarrow x - 6 = -3$$

$$\Rightarrow \boxed{x = +3}$$

Answer. 29 Let the oxidation number of Fe be x in Fe_3O_4. Then,

$$3x + 4(-2) = 0 \Rightarrow 3x - 8 = 0$$

$$\Rightarrow \boxed{x = \frac{+8}{3}}$$

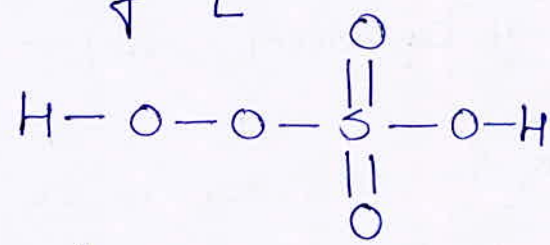
This fractional oxidation number is possible because Fe_3O_4 is a mixed oxide of $FeO.Fe_2O_3$. So, oxidation numbers of Fe in Fe_3O_4 is average value of $+2$ (O.N. in FeO) and $+3$ (O.N. in Fe_2O_3).

Answer.30 Since the net ionic charge on radical $C_2O_4^{2-}$ is -2; therefore the oxidation state of Mn must be $\underline{+2}$.

Answer.31 In the given reaction the oxidation number of Cl changes from +5 (in ClO_3^-) to −1 (in Cl^-). So, Cl atoms gains 6 electrons and thus 'X' should be $\underline{6e^-}$.

Answer.32 Same as solution of question 3.

Answer.33 Structure of H_2SO_5 :



Since the O atom in H_2SO_5 show peroxide linkage, hence evaluation of oxidation number should be made as:

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

$$\Rightarrow \boxed{x = +6}$$

Answer.34 Let the oxidation Number of I be x in KIO_4.

$$1 + x + 4(-2) = 0 \Rightarrow \boxed{x = +7}$$

In KI, O.N. of I = −1
In KI_3, O.N. of I = $-\frac{1}{3}$
In IF_5, O.N. of I = +5

Answer.35 Chemical formula of Hypochlorous acid is HClO. Let Oxidation number of Cl be x in HClO. Then,

$$1 + x - 2 = 0$$

$$\Rightarrow \boxed{x = +1}$$

Answer.31

Anion	Oxidation number
MnO_4^-	O.N. of Mn : +7
CrO_2Cl_2	O.N. of Cr : +6

So, both Mn and Cr are in their highest oxidation states.

Answer.32 $2H_3PO_4 + 3Ca(OH)_2 \longrightarrow Ca_3(PO_4)_2 + 6H_2O$

n-factor for acid $H_3PO_4 = 3$

So, equivalent weight of $H_3PO_4 = \frac{98}{3} = 32.66$

Answer.33 $2K\underset{+7}{Mn}O_4 + 16H\underset{-1}{Cl} \longrightarrow 2K\underset{-1}{Cl} + 2\overset{+2}{Mn}\underset{-1}{Cl_2} + 5\underset{0}{Cl_2} + 8H_2O$

Reduction

Oxidation

~~One mole of HCl requires~~

~~1 mole of KMnO4 gives the~~

1 mole of $KMnO_4$ requires 5 mole of electrons, thus
2 mole of $KMnO_4$ requires 10 mole of electrons.

1 mole of HCl gives 1 mole of electrons and thus, 16 mole of HCl gives 16 mole of electrons.

~~Thus, there is a~~ So, net transaction of ~~electrons~~ $\left(\frac{10}{16}\right)$ mole of electrons happen in the reaction.

$$\boxed{\text{So, Equivalent weight of HCl} = \frac{M}{\frac{\cancel{10}\,5}{\cancel{16}\,8}} = \frac{8M}{5}}$$

Answer.34

$$\underset{+4}{MnO_2} + 4\underset{-1}{HCl} \longrightarrow \underset{+2}{MnCl_2} + \underset{0}{Cl_2} + 2H_2O$$

Reduction

Oxidation

1 mole of MnO_2 requires 2 mole of electrons and 4 moles of HCl gives 4 mole of electrons. The net transaction $\frac{2}{4} = \frac{1}{2}$ mole of electrons happen in the reaction. So, $\boxed{\text{n-factor for HCl} = \frac{1}{2}}$

Answer.35 Since neutralisation is an acid-base reaction, thus we can consider the given compound to be working as ~~as~~ an acid because it contains 3 H^+ ions in it. Thus, n-factor for the compound will be 3. Thus,

$$\boxed{\text{Equivalent weight} = \frac{\text{Molecular weight}}{3}}$$

Level – 02

Answer. 1 $2[Fe(CN)_6]^{3-} + 2OH^- + \underset{-1}{H_2O_2} \longrightarrow 2[Fe(CN)_6]^{4-} + 2H_2O + \underset{0}{O_2}$

Oxidation

So, H_2O_2 is working as reducing agent ~~and~~ with n-factor = 2. Thus, $\boxed{E.W = \frac{34}{2} = 17}$

Answer. 2 $\underset{(O.N. = +4)}{PbO_2} \longrightarrow \underset{(O.N. = +2)}{Pb(NO_3)_2}$

So, PbO_2 is reduced.

Answer. 3 Since CO is a molecule with zero charge on it, therefore the oxidation state of metal in all the compounds is zero.

Answer. 4

		Oxidation state of P
Orthophosphorous acid	$\equiv H_3PO_3$	+3
Orthophosphoric acid	$\equiv H_3PO_4$	+5
Pyrophosphoric acid	$\equiv H_4P_2O_7$	+5
Metaphosphoric acid	$\equiv HPO_3$	+5

Answer. 5 Since the oxidation state of F is ~~also~~ always −1, therefore oxidation number of F in HOF is −1

Answer. 6 In reaction, $2Fe_2O_3 \cdot Cr_2O_3 + 4Na_2CO_3 + 3O_2 \longrightarrow 2Fe_2O_3 + 4Na_2CrO_4 + 4CO_2$

O.N. of Cr in $Fe_2O_3 \cdot Cr_2O_3 = +3$ $\quad (2x + 3(-2) = 0)$

O.N. of Cr in $Na_2CrO_4 = +6$ $\quad (2(+1) + x + 4(-2) = 0)$

Answer. 7 In $[Co(NH_3)_5Cl]Cl_2$, since NH_3 is a neutral molecule, therefore, oxidation state of N is −3.

In NH_2OH, $x + 2(1) + (-2) + (+1) = 0 \Rightarrow \boxed{x = -1}$

In Mg_3N_2, $3(+2) + 2x = 0 \Rightarrow \boxed{x = -3}$

In $(N_2H_5)_2SO_4$, to be neutral molecule the charge on N_2H_5 should be $\underline{+1}$ so that it can balance the -2 charge of SO_4^{2-}. Let the oxidation number of N in $(N_2H_5)^{+1}$ be x.

Then, $2x + 5(1) = 1 \Rightarrow \boxed{x = -2}$. So, option (3) is wrong.

$\boxed{\text{Answer. 8}}$ As we have seen in solution of question 4, only phosphorous acid has oxidation number of P to be $\underline{+3}$ and all others have oxidation number of P to be $\underline{+5}$.

$\boxed{\text{Answer. 9}}$ $[Ag(NH_3)_2]^+ + 2H^+ \longrightarrow Ag^+ + 2NH_4^+$

O.N. of Ag = +1   O.N. of Ag = +1   O.N. of N = -3

O.N. of N = -3

So, there is neither oxidation nor reduction.

$\boxed{\text{Answer. 10}}$ The net charge on $[Fe(H_2O)_5(NO)^+]$ should be $+2$ to balance -2 charge of SO_4^{2-}. The oxidation number on Fe be x, then,

$$\underset{\text{(For Fe)}}{x} + \underset{\text{(For } H_2O)}{0} + \underset{\text{(For } (NO)^+)}{1} = +2 \Rightarrow \boxed{x = +1}$$

$\boxed{\text{Answer. 11}}$ In reaction, $\underset{(\text{O.N. of Cr} = +6)}{CrO_4^{2-}} \longrightarrow \underset{(\text{O.N of Cr} = +6)}{Cr_2O_7^{2-}}$

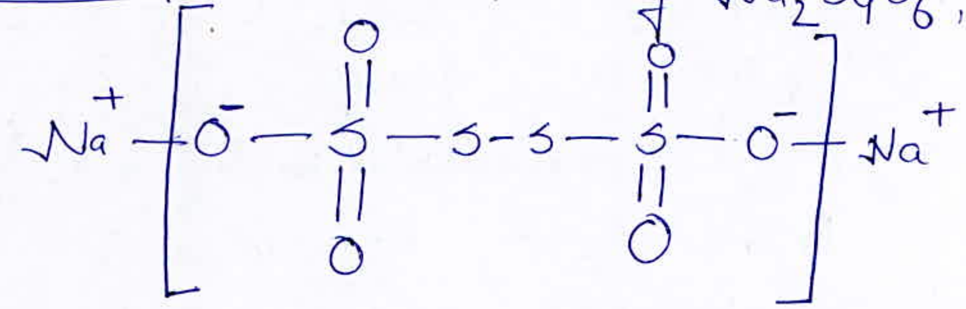
So, neither oxidation nor reduction.

$\boxed{\text{Answer. 12}}$ In any acidic medium,

~~$KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow$~~ $Mn^{+7} + 5e^- \longrightarrow Mn^{+2}$

So, oxidation number of Mn changes from $+7$ to $+2$.

Answer. 13 The structure of $Na_2S_4O_6$. (3)



Oxidation number of each S atom in S–S pure covalent bond is <u>zero</u>.
Oxidation number of each terminal S atom will be <u>+5</u> if we consider all sulphur Oxygen bonds to be completely ionic.

Answer. 14 $K_2Cr_2O_7 \longrightarrow Cr_2(SO_4)_3$
(O.N. of $Cr = +6$) (O.N. of $Cr = +3$)

Each Cr atom in $K_2Cr_2O_7$ undergoes change in oxidation state from +6 to +3. So,
n-factor for $K_2Cr_2O_7 = 2 \times (3) = 6$
Now, Equivalent weight (E) $= \frac{M}{\text{n-factor}} = \frac{M}{6}$
So, for $K_2Cr_2O_7$, $\boxed{M = 6E}$

Answer. 15 $Fe_3O_4 \longrightarrow Fe_2O_3$
(O.N. of $Fe = +\frac{8}{3}$) (O.N. of $Fe = +3$)

Each Fe in Fe_3O_4 undergoes change in oxidation state from $+\frac{8}{3}$ to +3. So, change in oxidation number per Fe atom in Fe_3O_4 is $\frac{1}{3}$. For one mole of Fe_3O_4, total change in oxidation number $= 3 \times \frac{1}{3} = 1$.
So, for Fe_3O_4, $E = \frac{M}{1} \Rightarrow \boxed{E = M}$

Answer.16 In reaction, $\underset{(O.N.\ of\ Ti = +3)}{2TiCl_3} \xrightarrow{\text{reduction}} \underset{(O.N.\ of\ Ti = +2)}{TiCl_2} + \underset{(O.N.\ of\ Ti = +4)}{TiCl_4}$

Oxidation

~~So, this~~ Since Ti is participating in both oxidation and reduction reactions. Therefore, the given reaction is an example of Disproportionation reaction.

Answer.17 The oxidation number decreases in the process of reduction.

Answer.18 $2\underset{(+3)}{Fe}Cl_3 + H_2\underset{(-2)}{S} \longrightarrow 2\underset{(+2)}{Fe}Cl_2 + 2HCl + \underset{(0)}{S}$

Since $FeCl_3$ is getting reduced (from +3 to +2) and H_2S is getting oxidised (from -2 to 0), hence $FeCl_3$ is working as an oxidant.

Answer.19 $3CuO + 2\underset{(-3)}{N}H_3 \longrightarrow 3Cu + \underset{(0)}{N_2} + 3H_2O$

So, in this reaction, the oxidation number of nitrogen is changing from −3 to 0.

Answer.20 $\underset{(+5)}{I}O_3^{-1} + H\underset{(-1)}{I} \longrightarrow H_2O + \underset{(0)}{I_2}$

increases

decreases

So oxidation number of I increases as well as decreases.

Answer.21 Let the oxidation number of Cu be x. Then,

$$+3 + 2(+2) + 3x + 7(-2) = 0$$

$$\Rightarrow \boxed{x = 7/3}$$

(5)

Answer. 22 $$\underset{(+6)}{WO_3} + 8CN^- + 2H_2\underset{(-2)}{O} \longrightarrow [\underset{(+4)}{W}(CN)_8]^{4-} + \frac{1}{2}\underset{(0)}{O_2} + 4OH^-$$

Oxidation

reduction

So, WO_3 is working as an oxidant.

Answer. 23 $$x\underset{+7}{Cl}O_4^- + y\underset{+4}{Mn}O_2 + zOH^- \longrightarrow x\underset{+1}{Cl}O^- + y\underset{+6}{Mn}O_4^{2-} + \frac{z}{2}H_2O$$

Reduction half: $ClO_4^- + 6e^- \longrightarrow ClO^- \longrightarrow (1)$

Oxidation half: $MnO_2 \longrightarrow MnO_4^{2-} + 2e^- \longrightarrow (2)$

To balance the number of electrons, multibly oxidation reaction (2) with 3, we get,

$$3MnO_2 \longrightarrow 3MnO_4^{2-} + 6e^- \longrightarrow (3)$$

Adding equation, (1) and (3), we get,

$$ClO_4^- + 3MnO_2 \longrightarrow ClO^- + 3MnO_4^{2-}$$

To balance charge on both sides, add OH^- on reactant side,

$$ClO_4^- + 3MnO_2 + 6OH^- \longrightarrow ClO^- + 3MnO_4^{2-}$$

Now, for balancing H and O atoms, add H_2O on product side,

$$ClO_4^- + 3MnO_2 + 6OH^- \longrightarrow ClO^- + 3MnO_4^{2-} + 3H_2O$$

Answer. 24 In acidic medium, the net reaction between $KMnO_4$ and FeC_2O_4 will be,

$$2MnO_4^- + H^+ + 5C_2O_4^{2-} \longrightarrow 2Mn^{2+} + 10CO_2 + H_2O$$

So, number of moles of $KMnO_4$ required will be $\left(\frac{2}{5}\right)$.

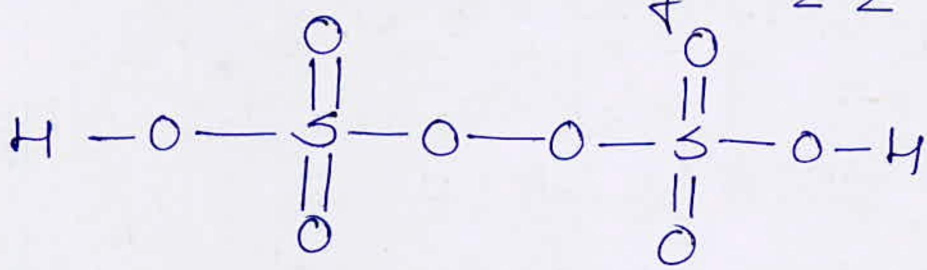
Answer. 25 In CCl_4, since oxidation state of Cl is -1 therefore, the oxidation state of C should be $\underline{+4}$.

Answer. 26

$$\underset{(+5)}{N_2}O_5 + H_2O \longrightarrow 2H\underset{(+5)}{N}O_3$$

So, the oxidation state of N doe not change.

Answer. 27 The structure of $H_2S_2O_8$ is



Considering all sulphur Oxygen bonds to be completely ionic, the oxidation number of each S atom would be $\underline{+6}$.

Answer. 28 Let the oxidation number of Cr in CrO_2Cl_2 is x.
Then, $x + 2(-2) + 2(-1) = 0 \Rightarrow \boxed{x = +6}$

Answer. 29

reduction

$$2Na_2\underset{(+2)}{S_2}O_3 + \underset{(0)}{I_2} \longrightarrow 2Na\underset{(-1)}{I} + Na_2\underset{\left(\frac{+5}{2}\right)}{S_4}O_6$$

Oxidation

So, since $Na_2S_2O_3$ is getting oxidised therefore it is acting as ~~an~~ a reducing agent.

Answer. 30 Oxidation state of oxygen is -1 ~~is~~ in peroxides.

Assertion and Reasoning Questions

$\boxed{\text{Answer. 1}}$ SO_2 and Cl_2 both are bleaching agents but SO_2 is both a reducing as well as an oxidisingg agent whereas Cl_2 is an oxidising agent.

$\boxed{\text{Answer. 2}}$ Flourine exists only in -1 oxidation state because it is the most electronegative element in the periodic table.

$\boxed{\text{Answer. 3}}$ $2HgCl_2 + SnCl_2 \longrightarrow SnCl_4 + Hg_2Cl_2$

So, $SnCl_2$ does not oxidises ~~no~~ $HgCl_2$ to Hg but itself get oxidised and reduces $HgCl_2$ to Hg_2Cl_2.

$\boxed{\text{Answer. 4}}$ $HClO_4$ is a stronger acid than $HClO_3$ because ClO_4^- ion is resonance stabilised but there is ~~no~~ less resonance stabilisaction in ClO_3^-.

$\boxed{\text{Answer. 5}}$

Oxidation

$$\underset{0}{Zn} + \underset{+2}{Cu}SO_4 \longrightarrow \underset{+2}{Zn}SO_4 + \underset{0}{Cu}$$

reduction

So, Zn is getting oxidised and thus it is acting as a reductant. The reason given is also a true explaination of the assertion.

$\boxed{\text{Answer. 6}}$ Let oxidation number of C be x. Then,

$$x + 2(+1) + (-2) = 0 \Rightarrow \boxed{x = 0}.$$

There is no relation between oxidation number of C being zero in HCHO and it being a covalent compound.

$\boxed{\text{Answer. 7}}$ Basic concepts about oxidation number.

$\boxed{\text{Answer. 8}}$ In H_2SO_4, let oxidation number of S be x. Then, $2(+1) + x + 4(-2) = 0 \Rightarrow \boxed{x = +6}$

Since S cannot increase its oxidation number ~~by~~ beyond +6, thus it cannot be oxidised further and therefore cannot work as an oxidising agent.

$\boxed{\text{Answer. 9}}$

$$\underset{0}{N_2} \longrightarrow \underset{-3}{NH_3}$$

So, one mole NH_3 can take 3 mole of electrons. Thus, Equivalent weight = $\dfrac{\text{Molecular weight}}{\text{no. of electron } e^- \text{ lost or gained}}$

$$\text{Equivalent weight of } NH_3 = \frac{17}{3}$$

One mole of N_2 can give $(2 \times 3) = 6$ mole of electrons. Thus,

$$\text{Equivalent weight of } N_2 = \frac{28}{6}$$

Previous Year's Questions

Answer. 1 (a) $H_2\underset{-1}{\underline{O}}_2 + O_3 \longrightarrow H_2\underset{-2}{\underline{O}} + 2O_2$

Reduction

(b) $H_2\underset{-1}{\underline{O}}_2 + Ag_2O \longrightarrow 2Ag + H_2\underset{-2}{\underline{O}} + \underset{0}{\underline{O}}_2$

Oxidation

Reduction

So, H_2O_2 is reducing in both (a) and (b).

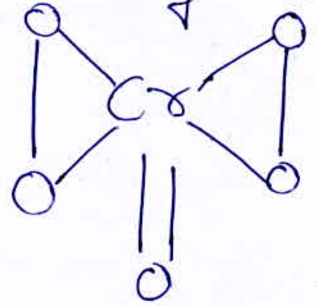
Answer. 3 In acidic medium, $KMnO_4$ undergoes reduction as

$$Mn^{+7} + 5e \longrightarrow \underline{Mn^{+2}}$$

Since $KMnO_4$ is getting reduced therefore we can say H_2O_2 is getting oxidised, so,

$$H_2O_2 \longrightarrow \underline{O_2}$$

Answer. 4 The structure of CrO_5 has two peroxide bonds



(A butterfly structure)

$$x + 4(-1) + 1(-2) = 0 \Rightarrow \boxed{x = +6}$$

Answer. 5 Reaction should be,

$$3\underset{0}{\underline{Cl}}_2 + 6NaOH \longrightarrow 5Na\underset{-1}{\underline{Cl}} + Na\underset{+5}{\underline{Cl}}O_3 + 3H_2O$$

So oxidation state of Cl changes from 0 to -1 and $+5$.

Answer. 6

Compound	Oxidation state of N
N_3H	$-\frac{1}{3}$ (maximum)
NH_2OH	-1
N_2H_4	-2
NH_3	-3 (minimum)

Answer. 7 On heating, $K\underset{+5}{\underline{Cl}}O_3 \longrightarrow K\underset{-1}{\underline{Cl}}$

$$H_2\underset{+6}{\underline{C_2}}O_4 \longrightarrow \underset{+4}{\underline{C}}O_2$$

$$H_2\underset{+6}{\underline{S}}O_4 \longrightarrow \underset{+4}{\underline{S}}O_2$$

So, Cl undergoes maximum change in oxidation number.

Answer. 8 $2\underline{S_2}O_3^{2-} \rightleftharpoons \underline{S_4}O_6^{2-}$

(O.N. of S = +2) (O.N. of S = $+\frac{5}{2}$)

Since oxidation number of S is increasing in this reaction, therefore it needs an oxidising agent.

Answer. 9

$$H_2O_2 + H_2O_2 \underset{\text{Reduction}}{\overset{\text{Oxidation}}{\rightleftharpoons}} 2H_2O + O_2$$

So, oxygen is both oxidised and reduced.

Answer. 10 In oxidation process, there is loss of electrons and thus number of electrons decreases.

Answer. 11 Since F is the most electronegative atom, therefore it has highest tendency to gain electron and thus most powerful oxidising agent.

Answer. 12

$$\overset{0}{H_2} + \overset{0}{Br_2} \longrightarrow 2\overset{+1}{H}\overset{-1}{Br}$$

Oxidation: H ($0 \to +1$); reduction: Br ($0 \to -1$)

Answer. 13

$$SO_2 + 2H_2\underset{(-2)}{S} \longrightarrow 3\underset{0}{S} + 2H_2O$$

Oxidation: S ($-2 \to 0$)

Thus, H_2S is getting oxidised.

Answer. 14 The net reaction of $KMnO_4$ is,

$$Mn^{+7} + 5e^- \longrightarrow Mn^{2+}$$

So, Equivalent weight $= \frac{M}{5}$

Answer. 15 In reaction, $Cl_2 + 2K\underset{-1}{I} \longrightarrow 2KCl + \underset{0}{I_2}$

Oxidation: I ($-1 \to 0$)

Answer. 16 Since ionisation energy of 'K' is least, therefore it has highest ~~reducing ag~~ tendency to lose electron and thus strongest reducing agent.

Answer. 17

$$\underset{0}{I_2} + 2\underset{+2}{S_2}O_3^{2-} \longrightarrow \underset{+5/2}{S_4}O_6^{2-} + 2\underset{-1}{I^-}$$

Oxidation: S ($+2 \to +\frac{5}{2}$); reduction: I ($0 \to -1$)

Iodine is reduced and Sulphur is oxidised.

Answer. 18 H_2O_2 can act as reducing agent and gives O_2 and it can also act as an oxidising agent and produce H_2O.

$$H_2O_2 \longrightarrow O_2 \quad \text{(Oxidation)}$$

$$H_2O_2 \longrightarrow H_2O \quad \text{(Reduction)}$$

Answer. 19 Since in $NaNO_2$, the oxidation state of of N is $\underline{+3}$ which is an intermediate oxidation state, therefore $NaNO_2$ can act as both oxidising and reducing agent.

Answer. 20

$$2\underset{+2}{\underline{Cu}}SO_4 + 4K\underset{-1}{\underline{I}} \longrightarrow \underset{+1}{\underline{Cu_2}}I_2 + 2K_2SO_4 + \underset{0}{\underline{I_2}}$$

Oxidation

Reduction

Answer. 21

$$H_2O + \underset{0}{\underline{Br_2}} \longrightarrow H\underset{-1}{\underline{Br}} + HO\underset{+1}{\underline{Br}}$$

Reduction

Oxidation

Answer. 22

$$Cu^{2+} + Ag \longrightarrow Cu + Ag^+$$

So, oxidation half : $Ag \longrightarrow Ag^+$

Answer. 23

$$2HI + H_2\underset{(+6)}{\underline{S}}O_4 \longrightarrow I_2 + \underset{(+4)}{\underline{S}}O_2 + 2H_2O$$

Reduction ~~Oxidising~~

So, H_2SO_4 is working as an oxidising agent.

Answer. 24

$$2H\underset{(+5)}{\underline{N}}O_3 + P_2O_5 \longrightarrow 2HPO_3 + \underset{(+5)}{\underline{N_2}}O_5$$

So, there is neither oxidation nor reduction.

Answer. 25

$$2K[\underset{(+1)}{\underline{Ag}}(CN)_2] + \underset{0}{\underline{Zn}} \longrightarrow 2\underset{(0)}{\underline{Ag}} + K_2[\underset{(+2)}{\underline{Zn}}(CN)_4]$$

Oxidation

reduction

So, this is a redox reaction.

Answer. 26

$$\underset{(+1)}{\underline{Ag_2}}O + H_2O_2 \longrightarrow 2\underset{0}{\underline{Ag}} + H_2O + O_2$$

reduction

So, H_2O_2 acts as reducing agent.

Answer. 27

Oxidation

$$4\underset{0}{P} + 3NaOH + 3H_2O \longrightarrow \underset{-3}{P}H_3 + 3NaH_2\underset{+1}{P}O_2$$

Reduction

Since P is getting both reduced and oxidised, so this is a disproportionation reaction.

Answer. 28 Since in CO_2, C is already present in its highest oxidation state of (+4), therefore CO_2 cannot act as reducing agent.

Answer. 29

reduction

$$H\underset{+1}{As}O_2 + Sn^{+2} \longrightarrow \underset{0}{As} + Sn^{4+} + H_2O$$

Oxidation

So, $HAsO_2$ is acting as an oxidising agent.

Answer. 30

$$2K\underset{(+7)}{Mn}O_4 + 16HCl \longrightarrow 5Cl_2 + 2\underset{(+2)}{Mn}Cl_2 + 2KCl + 8H_2O$$

reduction

So, the reduction product is $MnCl_2$.

Answer. 31

Oxidation

$$\underset{+2}{Ag^{2+}}(aq) + \underset{0}{Ag}(s) \rightleftharpoons 2\underset{+1}{Ag^{+}}(aq)$$

Reduction

This is an example of of comproportionation reaction. A comproportionationation reaction is a reaction between two reactants containing the same elements but different oxidation numbers to form a product. Comproportionation is the reverse reaction of a disproportionation reaction.

$\boxed{\text{Answer. 32}}$

$$\underset{0}{H_2} + Li \xrightarrow{} Li\underset{-1}{H}$$

reduction

~~So, b~~ since hydrogen is getting reduced ~~with~~ hence it is working as an oxidising agent.

Oxidation number

Answer. 1 The chemical formula of chromite ore is $FeCr_2O_4$. Since the charge on Cr_2O_4 ion is -2, therefore the oxidation numbers of Fe and Cr should be +2 and +3 respectively.

Answer. 2

Compound	Oxidation states of P
$H_4P_2O_5$	$4(+1)+2x+5(-2)=0 \Rightarrow \boxed{x=+3}$
$H_4P_2O_6$	$4(+1)+2x+6(-2)=0 \Rightarrow \boxed{x=+4}$
$H_4P_2O_7$	$4(+1)+2x+7(-2)=0 \Rightarrow \boxed{x=+5}$

Answer. 3

	Compound	Oxidation number of S
H_2SO_5	$\cancel{Na_2S_4O_6}$	$H-O-O-\underset{\underset{O}{\Vert}}{\overset{\overset{O}{\Vert}}{S}}-O-H \Rightarrow$ O.N. of S $= \underline{+6}$
H_2SO_3	$\cancel{H_2S_2O_7}$	$2x+x+3(-2)=0 \Rightarrow \boxed{x=+4}$
	SCl_2	$x+2(-1)=0 \Rightarrow \boxed{x=+2}$
	H_2S	$2(+1)+x=0 \Rightarrow \boxed{x=-2}$

Answer. 4 In $KMnO_4$, Mn has highest oxidation state of +7.

Answer. 5 Mn gives +7 oxidation state since its atomic number is 25.

Answer 6 In Cl_2O, the oxidation number of Cl should be $\underline{+1}$ because oxidation number of O is -2.

Answer 7 In HNO_3, let oxidation number of N be x. Then, $1 + x + 3(-2) = 0 \Rightarrow \boxed{x = +5}$

Answer 8 In methanal (HCHO) and methanoic acid (HCOOH), the oxidation numbers of C ~~is~~ are 0 and +2 respectively.

Answer 9 Let the oxidation state of S be x. Then,

$$2(+1) + x + 4(-2) = 0 \Rightarrow \boxed{x = +6}$$

Answer 10 Due to presence of NaOH, the medium is basic and thus the oxidation state of Mn changes from +2 to +6.

Answer 11 In S_8, oxidation number and covalency of S are 0 and 8 respectively.

Answer 12

Compound	Oxidation number
PO_4^{3-}	of $\underline{P}$: $x + 4(-2) = -3 \Rightarrow \boxed{x = +5}$
SO_4^{2-}	of $\underline{S}$: $x + 4(-2) = -2 \Rightarrow \boxed{x = +6}$
$Cr_2O_7^{2-}$	of $\underline{Cr}$: $2x + 7(-2) = -2 \Rightarrow \boxed{x = +6}$

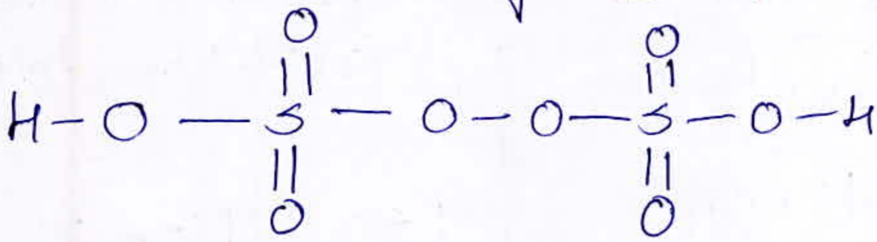
Answer 13 In $[Fe(H_2O)_6]Cl_3$, oxidation number of iron is $\underline{+3}$ because the net charge on complex ion is +3 and H_2O is a neutral molecule.

Answer 14 $Na_2O_2 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O_2$

Ion displacement reaction and thus no change in $\underline{O.N}$.

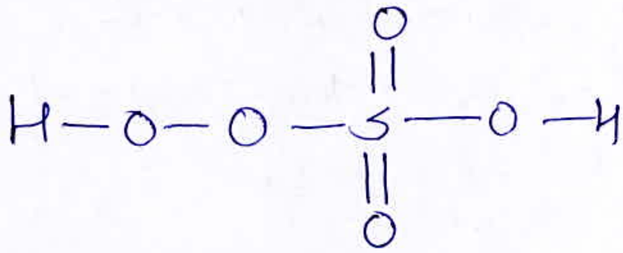
Answer. 15 The net charge on the complex ion $[Fe(H_2O)_5(NO)^+]$ should be +2. Therefore, the oxidation number of Fe should be $\underline{+1}$.

Answer. 16 The structure of $H_2S_2O_8$ is,



Considering all sulphur oxygen bonds to be completely ionic, the oxidation number of each S atom would be $\underline{+6}$.

Answer. 17 In peroxomonosulphuric acid (H_2SO_5),



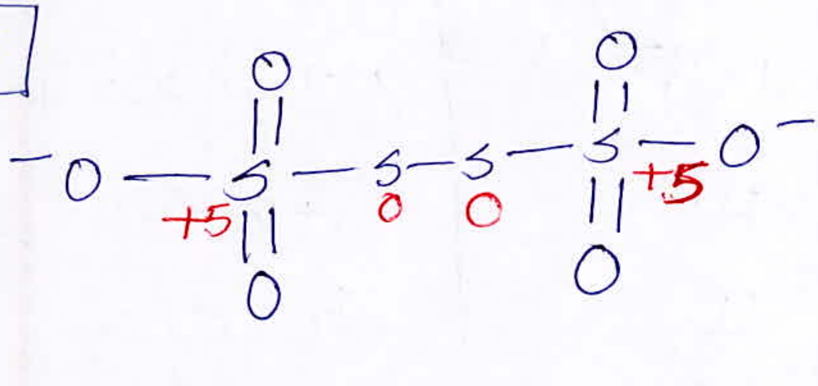
$$2 \times (+1) + x + 3(-2) + 2(-1) = 0$$

$$\Rightarrow \boxed{x = +6}$$

For peroxodisulphuric acid $(H_2S_2O_8)$,

$\boxed{x = +6} \longrightarrow$ As in answer. 16

Answer. 18



Answer.19 $Br_2 \longrightarrow BrO_3^-$

$\underline{0} \qquad \underline{+5}$

So, the oxidation number of Br changes from 0 to +5.

Answer.20

Compound	Oxidation state of I
HIO_4	$1 + x + 4(-2) = 0 \Rightarrow \boxed{x = +7}$
H_3IO_5	$3(+1) + x + 5(-2) = 0 \Rightarrow \boxed{x = +7}$
H_5IO_6	$5(+1) + x + 6(-2) = 0 \Rightarrow \cancel{7}\ \boxed{x = +7}$

Answer.21 Since oxidation number of F is always −1, therefore the oxidation number of <u>O</u> in F_2O is <u>+2</u>.

Answer.22 Since $K_2Cr_2O_7$ is an oxidising agent, therefore ~~SO₄~~ the oxidation number of S should change from +4 to +6.

Answer.23

Compound	Oxidation number of I
IO_3^-	$x + 3(-2) = -1 \Rightarrow \boxed{+5 = x}$
IO_4^-	$x + 4(-2) = -1 \Rightarrow \boxed{+7 = x}$
KI	$1 + x = 0 \Rightarrow \boxed{x = -1}$
I_2	$\boxed{x = 0}$

Answer. 24 Due to inert pair effect, lower oxidation states will be more stable, so the most characteristics oxidation states of lead and tin should be +2 and +4.

Answer. 25 The electronic configuration of Fe is $4s^2, 3d^6$. Due to stability of half filled orbitals, it can loose first 3 electrons easily and thus +3 is its most common oxidation state.

Answer. 26 Let the oxidation state of Cl in $KClO_4$ be x. Then,

$$1 + x + 4(-2) = 0 \Rightarrow \boxed{x = +7}$$

Answer. 27 Since maximum oxidation state possible for an element is +7, therefore the atom with electronic configuration $(n-1)d^5, ns^2$ will achieve highest oxidation state of +7.

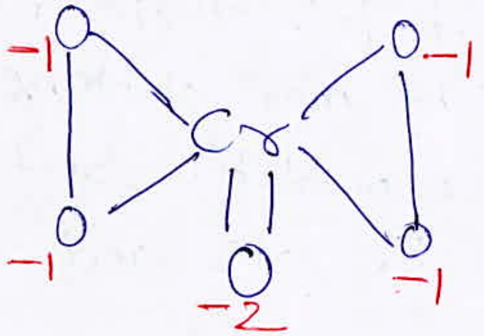
Answer. 28 Let the oxidation number of N be x. Then,

$$1 + x + 2(-2) = 0 \Rightarrow \boxed{x = +3}$$

Answer. 29 Since in case of CH_4, oxidation state of H is +1, thus oxidation state of C is -4, ~~thus~~ which is its lowest oxidation state.

Answer. 30 In NH_2OH, the oxidation state of N is -1 and not +1.

Answer. 31



Butterfly structure

Let the oxidation number of Cr to be $\underline{x}$. Then,

$$4(-1) + x + 1(-2) = 0 \Rightarrow \boxed{x = +6}$$

Answer. 32 Since the oxidation number of H is +1, therefore the oxidation number of N should be $\underline{-\frac{1}{3}}$ in N_3H.

Answer. 33 The net charge on (H_2PO_2) should be $\underline{-1}$ to balance the +2 charge on Ba. Let the oxidation number of P to be x. Then, $2(+1) + x + 2(-2) = -1$

$$\Rightarrow \boxed{x = +1}$$

Answer. 34 Pyrophosphoric acid $\equiv H_4P_2O_7$. Let oxidation number of P be x. Then,

$$4(+1) + 2x + 7(-2) = 0 \Rightarrow \boxed{x = +5}$$

Answer. 35 Let oxidation number of Xe be x. Then,

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

$$\Rightarrow \boxed{x = +4}$$

Answer. 36 To balance the charge of +1 of K, the net charge on $[Co(CO)_4]$ should be $\underline{-1}$. ~~Thus,~~ Since CO is a neutral molecule, therefore, the oxidation number of Co should be $\underline{-1}$.

Answer. 37 Let oxidation number of C be x. Then,

$$x + 2(+1) + 2(-1) = 0 \Rightarrow \boxed{x = 0}$$

Answer. 38

Radical	Oxidation number of S
SO_3^{2-}	$x + 3(-2) = -2 \Rightarrow \boxed{x = +4}$
$S_2O_4^{2-}$	$2x + 4(-2) = -2 \Rightarrow \boxed{x = +3}$
$S_2O_6^{2-}$	$2x + 6(-2) = -2 \Rightarrow \boxed{x = +5}$

So; $\boxed{S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}}$

Answer. 39 Since the element exhibits an oxidation state of +2, hence we can say that it should have <u>2</u> electrons in its outermost shell.

Answer. 40 Since there are 5 electrons in outermost shell, therefore ~~either~~ to achieve an octet it can take 3 electrons or to achieve a duplet it can loose all 5 electrons. Therefore, the oxidation state of N vary from <u>−3 to +5.</u>

Balancing of Chemical equations

Answer. 1 Reduction half: $Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{+3} \longrightarrow$ 1

Oxidation half: $2I^- \longrightarrow I_2 + 2e^- \longrightarrow$ 2

To balance the number of electrons, multiply equation 2 by 3, then,

$$6I^- \longrightarrow 3I_2 + 6e^- \longrightarrow 3$$

Adding equation 1 and 3, we get the net ionic equation as,

$$Cr_2O_7^{2-} + 6I^- \longrightarrow 2Cr^{+3} + \underline{3}I_2$$

So, 3 moles of I_2 will be liberated.

Answer. 2 $$x\underset{+7}{\underline{Mn}}O_4^- + y\underset{\cancel{+2}\ +3}{\underline{C}_2}O_4^{2-} + 3H^+ \longrightarrow \underset{+2}{\underline{Mn}^{+2}} + \underset{+4}{\underline{CO}_2} + H_2O$$

~~Reduction half: $Mn^{+7} \longrightarrow Mn^{2+}$~~

Reduction half: $Mn^{+7} \longrightarrow Mn^{+2} \longrightarrow$ 1

Oxidation half: $C_2O_4^{2-} \longrightarrow CO_2 \longrightarrow$ 2

Balancing of the atoms that undergo change in oxidation state. $C_2O_4^{2-} \longrightarrow 2CO_2$

Add the electrons lost or gained to each half equation

$$Mn^{+7} + 5e^- \longrightarrow Mn^{+2} \longrightarrow 3$$

$$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^- \longrightarrow 4$$

To balance the number of electrons, multiply equation ② ③ by 2 and equation ④ by 5, we get,

$$2Mn^{+7} + \cancel{10e^-} \longrightarrow 2Mn^{+2}$$

$$5C_2O_4^{2-} \longrightarrow 5CO_2 + \cancel{10e^-}$$

<u>Add</u>: $2Mn^{+7} + 5C_2O_4^{2-} \longrightarrow 2Mn^{+2} + 5CO_2$

<u>Main equation</u>: $2MnO_4^- + 5C_2O_4^{2-} \longrightarrow 2Mn^{2+} + 5CO_2$

To balance charge on both sides, use H^+, then,

$$\boxed{2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 5CO_2}$$

So, the values of x, y, z are <u>2</u>, <u>5</u> and <u>16</u> respectively.

$\boxed{\text{Answer. 3}}$ $x\underset{-1}{I^-} + y\underset{+5}{I}O_3^- + zH^+ \longrightarrow \underset{0}{I_2} + H_2O$

Reduction half: $IO_3^- + 5e^- \longrightarrow I_2 \longrightarrow$ ①.

Oxidation half: $2I^- \longrightarrow I_2 + 2e^- \longrightarrow$ ②

To balance the number of electrons, multiply equation ② by 5, we get

$$5I^- \longrightarrow 5I_2 + 5e^- \longrightarrow ③$$

Adding ① and ③, we get,

$$5I^- + IO_3^- \longrightarrow 5I_2$$

To balance charge, use H^+, then,

$$5I^- + IO_3^- + 6H^+ \longrightarrow 5I_2$$

So, $x = 5$, $y = 1$ and $z = 6$.

Answer. 4 $x\underset{+4}{U}O^{2+} + \underset{+6}{Cr_2}O_7^{2-} + yH^+ \longrightarrow a\underset{+6}{U}O_2^{2+} + z\underset{+3}{Cr}^{3+} + bH_2O$

Reduction half: $Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+} \longrightarrow (1)$

Oxidation half: $UO^{2+} \longrightarrow UO_2^{2+} + 2e^- \longrightarrow (2)$

For balancing electrons, multiplying equation (2) by $\underline{3}$,

$$3UO^{2+} \longrightarrow 3UO_2^{2+} + 6e^- \longrightarrow (3)$$

Adding (1) and (3), we get,

$$3UO^{2+} + Cr_2O_7^{2-} \longrightarrow 3UO_2^{2+} + 2Cr^{3+}$$

For balance of charge, add H^+ on reactant side,

$$3UO^{2+} + Cr_2O_7^{2-} + 8H^+ \longrightarrow 3UO_2^{2+} + 2Cr^{3+}$$

So, $\underline{x = 3,\ y = 8 \text{ and } z = 2}$.

Answer. 5 $xK\underset{+7}{\underline{Mn}}O_4 + \underset{-3}{\underline{N}}H_3 \longrightarrow yK\underset{+5}{\underline{N}}O_3 + \underset{+4}{\underline{Mn}}O_2 + KOH + H_2O$

Reduction half: $Mn^{+7} + 3e^- \longrightarrow Mn^{+4} \longrightarrow (1)$

Oxidation half: $\underline{N}H_3 \longrightarrow NO_3^- + 8e^- \longrightarrow (2)$

For balancing of electrons multibly equation (1) by 8 and equation (2) by 3, we get

$$8Mn^{+7} + \cancel{24e^-} \longrightarrow 7Mn^{+4}$$

$$3NH_3 \longrightarrow 3NO_3^- + \cancel{24e^-}$$

$$8Mn^{+7} + 3NH_3 \longrightarrow 7Mn^{+4} + 3NO_3^-$$

Substituting to main equation, we get,

$$\underset{x}{\underline{8}}KMnO_4 + 3NH_3 \longrightarrow 7MnO_2 + \underset{y}{\underline{3}}KNO_3$$

Answer. 6 $\underset{0}{Zn} + \underset{+5}{N}O_3^- \longrightarrow \underset{+2}{Zn}^{+2} + \underset{-3}{N}H_4^+$

Reduction half: $NO_3^- + 8e^- \longrightarrow NH_4^+ \longrightarrow (1)$

Oxidation half: $Zn \longrightarrow Zn^{+2} + 2e \longrightarrow (2)$

For balancing of electrons, multiplying equation (2) by 4, we get,

$$4Zn \longrightarrow 4Zn^{+2} + 8e^- \longrightarrow (3)$$

Adding (1) and (3), we get,

$$4Zn + NO_3^- \longrightarrow 4Zn^{+2} + NH_4^+$$

To balance charge, add OH^- on reactant side,

$$4Zn + NO_3^- + 10OH^- \longrightarrow 4Zn^{+2} + NH_4^+$$

So, $x = 4$, $y = 1$ and $z = 10$

Answer. 7 In alkaline medium, $Mn^{+7} + e^- \longrightarrow Mn^{+6}$

The equivalent mass of $KMnO_4$ in strong alkaline medium is its molar mass itself.

Answer. 8

$$\underset{+5}{N}O_3^- \longrightarrow \underset{-2}{N_2}H_4$$

So, the number of electrons involved in the reduction is 7.

Answer. 9 Reduction half: $IO_3^- + 5e^- \longrightarrow I_2$

Oxidation half: $SO_2 \longrightarrow SO_4^{2-} + 2e^-$

For balancing the number of electrons, we have to multiply oxidation reaction by 5, thus, the coefficient of SO_2 in main reaction will be 5.

Answer. 10 $\underset{+5}{Bi}O_3^- + 6H^+ + 2e^- \longrightarrow \underset{+3}{Bi^{3+}} + 3H_2O$

Answer. 11 Oxidation half: $H_2O_2 \longrightarrow O_2 + 2e^-$

Reduction half: $ClO_2 + 5e^- \longrightarrow Cl^-$

To balance the number of electrons, multibly oxidation reaction with 5, then we have,

$$5H_2O_2 \longrightarrow 5O_2 + 10\cancel{e^-}$$

$$10\cancel{e^-} + 2ClO_2 \longrightarrow 2Cl^-$$

$$5H_2O_2 + 2ClO_2 \longrightarrow 5O_2 + 2Cl^-$$

so, one mole of ClO_2 oxidises $\left(\frac{5}{2} = 2.5\right)$ moles of H_2O_2.

Answer. 12 $Cu + 4HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

Answer. 13 $H_2\underset{+6}{S}O_4 + xH\underset{-1}{I} \longrightarrow H_2\underset{-2}{S} + y\underset{0}{I_2} + zH_2O$

Reduction half: $H_2SO_4 + 8e^- \longrightarrow H_2S \longrightarrow (1)$

Oxidation half:

$$I^- \longrightarrow \frac{1}{2}I_2 + e^- \longrightarrow (2)$$

To balance number of electrons, multibly equation (2) by 8, we get, $8I^- \longrightarrow 4I_2 + 8e^- \longrightarrow (3)$

Adding (1) and (3), we get,

$$H_2SO_4 + 8HI \longrightarrow H_2S + 4I_2$$

To balance H and O, add H_2O molecules on product side,

$$H_2SO_4 + 8HI \longrightarrow H_2S + 4I_2 + 4H_2O$$

So, $x = 8$, $y = 4$ and $z = 4$.