<H1>Additional Objective Questions
<H2>Single Correct Choice Type

1.(A) In H_2SO_4, sulphur has +6 oxidation state in SO_2 = +4, $Na_2S_2O_3$ = +2.

2.(D) In carbonyl compounds metal has zero oxidation state, as CO is neutral and has a zero state.

3.(C) Increase in oxidation number is oxidation

(i) $\overset{+3}{C_2}O_4^{2-} \rightarrow 2\overset{+4}{C}O_2$ oxidation

(ii) $\overset{+6}{S}O_4^{2-} \rightarrow \overset{+4}{S}O_3^{2-}$ reduction

(iii) $\overset{+6}{Mn}O_4^{2-} \rightarrow \overset{+7}{Mn}O_4^{-}$ oxidation

(iv) $Fe^{3+} \rightarrow Fe^{2+}$ reduction

4.(D) In the given reaction, it can be predicted that, $Na_2S_2O_3$ is a reducing agent and iodine is an oxidizing agent. Since, one is reduced and other is oxidized, therefore, this reaction is an example of intermolecular redox reaction. Due to the presence of S–S linkage in $Na_2S_4O_6$, the average oxidation state of S is +5/2.

5.(B) $2NO_3^- + 7e^- + 10H^+ \rightarrow N_2H_4 + 3H_2O$

6.(D) $SO_4^{2-} \Rightarrow +6 + x(-2) = -2 \Rightarrow x = 4.$

7.(C) The reaction involved is

Reduction

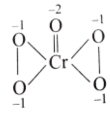
$$\overset{0}{Cl_2} + OH^- \rightarrow \overset{-1}{Cl^-} + \overset{+5}{Cl}O_3^- + H_2O$$

Oxidation

8.(B) For the given reaction,

$$\overset{+6}{Cr}O_5 \rightarrow \overset{+3}{Cr_2}(SO_4)_3$$

The structure of CrO_5 is



The number of moles of electron required = 3

9.(B) In acidic medium,

$$MnO_1^- \rightarrow Mn^{+2} \qquad n\text{-factor} = 5/\text{mol}$$

$$Cr_2O_7^{2-} \rightarrow 2Cr^{+3} \qquad n\text{-factor} = 6/\text{mol}$$

Hence, the amount of Fe(II) oxidized is more with $Cr_2O_7^{2-}$.

10.(D) In the reaction

$$I^{5+} + 6e^- \rightarrow I^-$$
$$S^{6+} + 2e^- \rightarrow S^{4+}$$

Therefore,

$$M_{equiv} NaHSO_3 = M_{equiv} NaIO_3$$
$$N_1V_1(NaHSO_3) = N_2V_2(NaIO_3)$$
$$\frac{w}{104/2} \times 100 = \frac{0.66}{198/6} \times 100$$
$$w = 1.04 \text{ g}$$

11.(C) 1 mol each of Hg^{2+} and I^- will form 1/4 = 0.25 mol.

12.(A) $Cr_2O_7^{2-} + 3Sn^{2+} + 14H^+ \rightarrow 3Sn^{4+} + 2Cr^{3+} + 7H_2O$

3 mol of Sn^{2+} react with 1 mol of $Cr_2O_7^{2-}$.

So, 1 mol of Sn^{2+} will react with (1/3) mole of $Cr_2O_7^{2-}$.

13.(D) The reaction involved is

$$3H_3\overset{+1}{P}O_2 \rightleftharpoons \overset{-3}{P}H_3 + 2H_3\overset{+3}{P}O_3$$

The number of electrons transferred in oxidation half reaction is 4 and in the reduction half reaction is 2.

Hence, the effective number of electrons transferred = 4, and total number of moles involved = 3.

Therefore, Equivalent weight = $\frac{\text{Molecular weight}}{n\text{-factor}} = \frac{M}{4/3} = \frac{3M}{4}$

14.(D) Let x be the oxidation state of Cr. Then $x + 2(-2) + 2(-1) = 0 \Rightarrow x = +6$. Similarly for the others the oxidation states are $\overset{+7}{Mn}O_4^-$, $\overset{+3}{Cr}(CN)_6^{3-}$, $\overset{+4}{Ni}F_6^{2-}$.

15.(B) The reaction involved is $5As_2O_3 + 4MnO_4^- + 9H_2O + 12H^+ \rightarrow 10HAsO_4^{3-} + 4Mn^{2+}$

5 mol of As_2O_3 reacts with 4 mol of MnO_4^-

Therefore, $0.46/198$ mol of As_2O_3 will reacts with $\frac{4}{5} \times \frac{0.46}{198} = 0.00186$ mol of MnO_4^-

Hence, molarity of $KMnO_4$ = $\frac{0.00186}{25} \times 1000 = 0.074$ M

16.(D) Milliequivalent of hypo = Milliequivalent of Cu^{2+}

= 25 mass equivalents of Cu^{+2} (n = 1) = 25 mol of Cu^{2+}

$$\frac{m \text{ mol of } \begin{matrix}\text{COOH}\\ | \\ \text{COOH}\end{matrix}}{m \text{ mol of } KMnO_4} = \frac{5}{2}$$

Therefore, m mol of oxalic acid $= \frac{5}{2} \times 10 = 25$

$$\frac{m \text{ mol of } Cu^{2+}}{m \text{ mol of oxalic acid}} = \frac{25}{25} = 1 \Rightarrow \frac{\text{Mole of } Cu^{2+}}{\text{Mole of oxalic acid}} = 1:1$$

17.(D) Volume of HCl neutralized by NaOH = (Caustic soda) = V_1

$$\text{(HCl) } N_1V_1 = N_2V_2 \text{ (NaOH)}$$

$$0.1 \times V_1 = 0.2 \times 30 \Rightarrow V_1 = 60 \text{ mL}$$

So, the remaining 40 mL 0.1 N HCl is now neutralized by 0.25 N KOH

$$\text{(HCl) } N_1V_1 = N_2V_2 \text{ (KOH)}$$

$$0.1 \times 40 = 0.25 \times V_2 \Rightarrow V_2 = 16 \text{ mL}$$

18.(A) Moles of I_2 formed = 3 × moles of KIO_3 $= 3 \times \frac{0.57}{214}$

Mass equivalent of I_2 formed (n-factor = 2) $= 3 \times \frac{0.57}{214} \times 2 \times 1000$

Let the volume of $Na_2S_2O_3$ consumed be V mL; then

$$(V \times 0.1) \times 8 = \frac{3 \times 0.57}{214} \times 2000 \Rightarrow V = 19.97 \text{ mL}$$

19.(C) For the given reaction

$$3\overset{+1}{Cl}O^-(aq) \rightarrow \overset{+5}{Cl}O_3^-(aq) + 2\overset{-1}{Cl^-}(aq)$$

ClO^- get reduced as well as oxidized, therefore the reaction is an example of disproportionation reaction.

20.(D) Disproportionation is the process in which an element undergoes both oxidation and reduction in a reaction. Moreover, an element can show a maximum oxidation state equal to its "group number" and a minimum oxidation state equal to its "group number 8". Among the oxyacids of chlorine given, in $HClO_4$ chlorine is in its highest oxidation state (+7). So, it cannot undergo further oxidation, and hence, $HClO_4$ cannot undergo disproportionation.

21.(A) Less positive charge, more basic oxide.

$$\overset{+2}{M}nO, \overset{+3}{M}n_2O_3, \overset{+4}{M}nO_2, \overset{+7}{M}n_2O_7$$

22.(B) The oxidation states are as follows:

$N^1 = 0$ $N^2 = 0$ $N^3 = -1$(as it is attached to H)

23.(A) Among the given oxidizing agents, ozone is the strongest oxidizing agent because on decomposition it forms atomic oxygen, which is highly energetic.

$$O_3 \rightarrow O_2 + O$$

24.(C) The reaction involved is

$$K\overset{+5}{Cl}O_3 + 4e^- \rightarrow KO\overset{+1}{Cl}^- + O_2$$

25.(A) $CaOCl_2$ has two chlorine atoms having different oxidation states.
Ca^{2+} $(OCl)^-$ Cl^- where $(OCl)^-$ is the hypochloride ion and Cl^- chloride ion. In the hypochoride ion, chlorine is in +1 state, while in chloride ion, it is in –1 state.

26.(D) The blue color appears due to *d–d* transition that takes place in copper when aqueous NH_3 (ligand) is added to the aqueous copper sulphate solution. The reaction involved is

$$\overset{2+}{Cu}\overset{2-}{S}O_4 + 4\overset{-3}{N}\overset{+1}{H}_3 \rightarrow [\overset{+2}{Cu}(\overset{-3}{N}\overset{+1}{H}_3)_4]SO_4$$

27.(C) The maximum oxidation state possible for Br = +7. In $KBrO_4$, Br exists in +7 oxidation state.

28.(B) Balanced equations for producing NO and NO_2, respectively, are

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O \quad (1)$$

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O \quad (2)$$

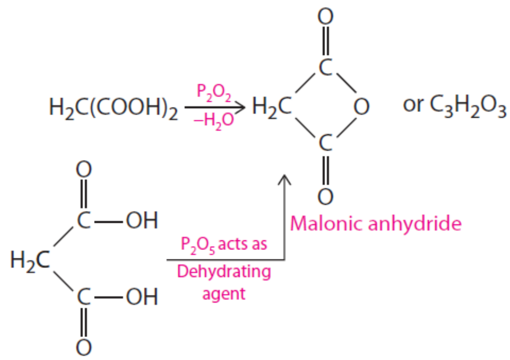
Adding Eqs. (1) and (2), we have

$$4Cu + 12HNO_3 \rightarrow 4Cu(NO_3)_2 + 2NO_2 + 2NO + 6H_2O$$

$$2Cu + 6HNO_3 \rightarrow 2Cu(NO_3)_2 + NO_2 + NO + 3H_2O$$

Thus, coefficients *x* and *y* of Cu and HNO_3 are 2 and 6, respectively.

29.(C) The reaction is



Malonic acid

Suppose oxidation state of carbon in $C_3H_2O_3$ = x

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

$\Rightarrow$ $$3x + 2 - 6 = 0$$

$\Rightarrow$ $$3x - 4 = 0$$

$\Rightarrow$ $$x = +4/3$$

30.(A) The oxidation states are

$$\underset{\text{Sodium thiosulphate}}{Na_2\overset{+2}{S_2}O_3} \rightarrow \underset{\text{Sodium tetrathionate}}{Na_2\overset{+2.5}{S_4}O_6}$$

Change in oxidation number of sulphur = 1/2.

31.(D) The reaction involved is

$$\underset{\text{Potassium dichromate}}{K_2\overset{+6}{Cr_2}O_7} \xrightarrow[\text{medium}]{\text{Alkaline}} \underset{\text{Potassium chromate}}{K_2\overset{+6}{Cr}O_4}$$

No change in the oxidation state of chromium takes place.

32.(C) Suppose oxidation number of C = x, then $x + 2 - 2 = 0 \Rightarrow x = 0$

33.(C) The reaction involved is

$$\overset{+6}{Cr_2}O_7^{2-} + \overset{+4}{S}O_2 \longrightarrow 2\overset{3+}{Cr} + \overset{+6}{S}O_4^{2-}$$

O.N + 4 + 0 + 6

34.(A) The oxidation state of S in its anions follows the order: $\overset{+3}{S_2}O_4^{2-} < \overset{+4}{S}O_3^{2-} < \overset{+5}{S_2}O_6^{2-}$

35.(A) In this compound K is present in the highest possible oxidation state.

$$\overset{+1}{K}\overset{-2}{S}\overset{+4}{C}\overset{-3}{N}$$

36.(D)

Element	%	Atomic mass	Relative no. of moles	Simple ratio of moles
Xe	53.5	131.3	53.5/131.3 = 0.41	1
F	46.5	19	46.5/19 = 2.5	6

Formula of the compound = XeF_6
Hence, oxidation number of Xe = +6.

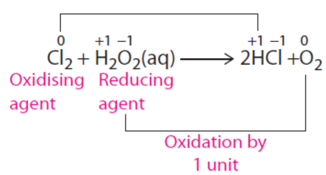
37.(A) $\overset{+6}{Cr}\overset{2-}{O_4} \rightarrow \overset{+6}{Cr_2}\overset{2-}{O_7}$
No change in oxidation number.

38.(D) Disproportionation is a type of reaction in which the same element is oxidized and reduced as well.

$$\overset{0}{P_4} + 2NaOH \rightarrow 3NaH_2\overset{+1}{P}O_2 + \overset{-3}{P}H_3$$

In the above reaction, P is oxidized to PH_3 and reduced to NaH_2PO_2.

39.(B) The reaction involved is



40.(C) The increasing order of the oxidation state of the nitrogen is: $\overset{-3}{N}H_4^+ < \overset{-2}{N_2}H_4 < \overset{-1}{N}H_2OH < \overset{+1}{N_2}O$.

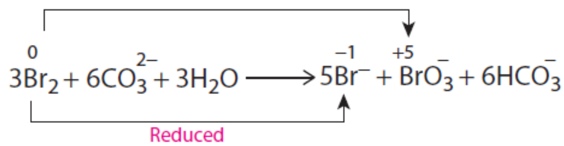
41.(C) K is at topmost position in the electrochemical series in comparison with other elements (Mg, Na, Br_2). Therefore, it will be the most strongly reducing substance.

42.(C) The compound in which an element is present in its highest or lowest oxidation state cannot undergo disproportionation reaction. In H_5IO_6, the oxidation number of I is +7 (its highest oxidation state).

43. In the reaction

$$3Br_2 + 6CO_3^{2-} + 3H_2O \rightarrow 5Br^- + BrO_3^- + 6HCO_3^-$$

43.(D) The reaction involved is



44.(D) The reaction involved is

$$CrO_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

$$[Sn^{2+} \rightarrow Sn^{4+} + 2e^-] \times 3$$

$$\underline{CrO_7^{2-} + 3Sn^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 3Sn^{4+} + 7H_2O}$$

45.(C) Total oxidation of two nitrogen atoms in N_2H_4 is –4. Since it loses 25 mol electrons, the total oxidation number of two N atoms in Y increases by

$$\overset{-2}{N_2}H_4 \quad \rightarrow \quad N_2^x \quad + \quad ne^-$$

$$2.5\,\text{mol} \qquad\qquad 25\,\text{mol}$$

$$1\,\text{mol} \qquad\qquad 10\,\text{mol}$$

In hydrazine, the oxidation number of N is $2x + 4\ (+1) = 0$ or $x = -2$

Change in oxidation number per N atom = number of electrons lost

Hence, $2(x + 2) = 10 \Rightarrow x = +3$.

46.(D) This is because fluorine is the most electronegative element in the periodic table and its oxidation state is always –1.

47.(B) Suppose oxidation number of sulphur = x

$$2 \times 1 + 4 \times (x) + 6 \times (-2) = 0 \Rightarrow 2 + 4x - 12 = 0$$

$$4x - 10 = 0 \Rightarrow 4x = 10 \Rightarrow x = +2.5$$

48.(C) In F_2O, O is in the +2 state.

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

49.(B) The reactions involved are

$$MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{+2} + 4H_2O$$

$$[Fe^{2+} \rightarrow Fe^{3+} + e^-] \times 5$$

$$\overline{MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{3+} + 4H_2O}$$

Hence,

5 mol of $FeSO_4$ are completely oxidized by $KMnO_4$ = 1 mol

15 mol of $FeSO_4$ are completely oxidized by $KMnO_4$ = $\frac{1\times15}{5}=3\,\text{mol}$

50.(C) The reactions are as follows:

$$\text{Acidic medium}: \overset{+7}{\text{M}}\text{nO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{+2} + 4\text{H}_2\text{O}$$

$$\text{Basic medium}: \overset{+7}{\text{M}}\text{nO}_4^- + e^- \rightarrow \overset{+6}{\text{M}}\text{nO}_4^{2-}$$

$$\text{Neutral medium}: \overset{+7}{\text{M}}\text{nO}_4^- + 2\text{H}_2\text{O} + 3e^- \rightarrow \overset{+4}{\text{M}}\text{nO}_2 + \overset{-}{\text{O}}\text{H}$$

<H2>Multiple Choice Questions

1.(B, C) A reducing agent in a redox reaction undergoes an increase in the oxidation number with loss of electrons.

2.(B, C, D) For the given reaction

$$2\overset{+1}{\text{Cu}}_2\overset{-2}{\text{O}}(\text{s}) + \overset{+1}{\text{Cu}}_2\overset{-2}{\text{S}}(\text{s}) \rightarrow 6\overset{0}{\text{Cu}}(\text{s}) + \overset{+4}{\text{S}}\overset{-2}{\text{O}}_2(\text{g})$$

Both Cu_2O and Cu_2S are reduced to Cu. Also, Cu_2O acts as the oxidant while Cu_2S acts as the reductant as it gets oxidized to SO_2.

3.(B, C, D) Except F_2 (which shows oxidation state of –1 only), all the other substances show both +ve and –ve oxidation states and hence undergo disproportionation.

4.(A, C) Sn and Cu can undergo disproportionation reaction.

For Cu:

$$\begin{aligned} &\text{Cu}^+ \rightarrow \text{Cu}^{2+} + e^- \\ &\underline{\text{Cu}^+ + e^- \rightarrow \text{Cu}} \\ \text{Overall}\quad &\underline{2\text{Cu}^+ \rightarrow \text{Cu}^{2+} + \text{Cu}} \end{aligned}$$

For Sn:

$$\begin{aligned} &\text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + 2e^- \\ &\underline{\text{Sn}^{2+} + 2e^- \rightarrow \text{Sn}} \\ \text{Overall}\quad &\underline{2\text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + \text{Sn}} \end{aligned}$$

5.(B, C, D) Statement (A) is not correct. For example, during neutralization reactions and precipitation reactions oxidation number does not change. Hence, these reactions (collectively known as metathesis reaction) are not redox reactions.

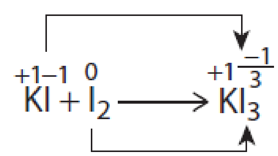
6.(B, C, D) In the given acid pairs, the one in which the central atom is present in its higher oxidation state is stronger acid than the other.

$$H_3\overset{+5}{A}sO_4 > H_3\overset{+3}{A}sO_3 \qquad H\overset{+7}{Cl}O_4 > H\overset{+5}{Cl}O_3 \qquad H\overset{+5}{N}O_3 > H\overset{+3}{N}O_2$$

7.(A, C, D) Suppose oxidation state of oxygen = x

B_4O_{10}	CsO_2	KO_3
$4 \times 1 + 10(x) = 0$	$+1 \times 2(x) = 0$	$+1 \times 3(x) = 0$
$4 + 10x = 0$	$1 \times 2x = 0$	$1 = -3x$
$4 = -10x$	$1 = -2x$	$1 = -3x$
$x = -\frac{4}{10}$	$x = -\frac{1}{2}$	$x = -\frac{1}{3}$

8.(A, B, C) The reaction involved is



9.(A, D) Because Cd is less reactive than Al and Ag is more reactive than Al.

10.(C, D) Both (C) and (D) are disproportionation reactions (redox). Redox reactions are those in which both reduction and oxidation occur simultaneously.

11.(A, B, C) For the given reaction

$$2\overset{+7}{Mn}O_4^- + 4H^+ + \overset{0}{Br_2} \rightarrow 2\overset{+2}{Mn}{}^{+2} + 2\overset{+5}{Br}O_3^- + 2H_2O$$

n-factor for $MnO_4^- = 5$ (as +7 to +2)

Equivalent weight = $M_x/5$

n-factor for Br_2 = 10 (as 0 to +5). Therefore, for two atoms = 5 × 2 = 10

Equivalent weight = $M_y/10$

Ratio of *n*-factor $\overset{+2}{Mn}{}^{2+} \rightarrow \overset{+7}{Mn}O_4^-$

$3\overset{+5}{Br}O_3^- \rightarrow \overset{0}{Br_2} = 1 : 1$

12.(A, B, C) Here, we estimate the concentration of Fe^{2+} ions; so, the options A, B, and C are the correct reasons. By rinsing the flask with solution of Fe^{2+}, the concentration of Fe^{2+} in the flask increases. If the burette is rinsed with $KMnO_4$, the concentration of KMnO4 in the burette reduces. Hence more volume of $KMnO_4$ is required for titration. Lastly, when the last drop of Fe^{2+} solution was blown from the pipette into the flask, the concentration of Fe^{2+} obtained is higher than the actual one. So, the options a, b, and c bring out discrepancies in the titration.

13.(A, C, D) NaOH and $NaHCO_3$ in the solution react as

$$\underset{40\text{ g}}{NaOH} + \underset{84\text{ g}}{NaHCO_3} \rightarrow Na_2CO_3 + H_2O$$

Now, number of moles of NaOH = $\frac{40}{40} = 1$ mol and of $NaHCO_3 = \frac{84}{84} = 1$ mol

After the reaction, the solution will have 2 mol of Na_2CO_3 in 1 L.

(i) For phenolphthalein,

Milliequivalent of Na_2CO_3 = Milliequivalent of HCl

$$20 \times 20 \times 1 = V \times 1 \Rightarrow V = 40 \text{ mL}$$

(ii) For methyl orange after the first end point

Milliequivalent of $NaHCO_3$ = Milliequivalent of Na_2CO_3 = Milliequivalent of HCl

$$2 \times 20 \times 1 = 1 \times V \Rightarrow V = 40 \text{ mL}$$

(iii) For methyl orange if used from the very beginning:

Milliequivalent Na_2CO_3 + Milliequivalent of $NaHCO_3$ = Milliequivalent of HCl

2 × Milliequivalentof Na_2CO_3 = Milliequivalent of HCl

$$2 \times 20 \times 2 = 1 \times V \Rightarrow V = 80 \text{ mL}$$

14.(A, B, C, D) From the reaction

$$Ba(\overset{+7}{M}nO_4)_2 \rightarrow 2\overset{+2}{M}n^{2+} \ (n\text{-factor} = 10)$$

Milliequivalent of $Ba(MnO_4)_2$ in 150 mL $\Rightarrow 150 \times 10 \times \frac{1}{10} = 150$ millieq.

Milliequivalent of 1M $Fe^{2+} \rightarrow Fe^{++}$ = 150 × 1 = 150m.Eq

Milliequivalent t of 50 mL 1M $FeC_2O_4 \rightarrow Fe^{3+} + 2CO_2$ = 50 × 3 = 150 m.Eq

Milliequivalent of 75 mL 1M $C_2O_4^{2-} \rightarrow 2CO_2 = 75 \times 2 = 150$ m.Eq

Milliequivalent of 25 mL 1M $C_2O_7^{2-} \rightarrow 2Cr^{3+} = 25 \times 6 = 150$ m.Eq

15.(A, B, C) In acidic medium

$$\overset{+7}{M}nO_4^- + 8H^+ + 5e^- \rightarrow \overset{+2}{M}n^{+2} + 4H_2O$$

Total number of electrons gained = 5

Total change in oxidation number of Mn = 7 – 2 = 5

Hence, equivalents of MnO_4^- = 0.1 × 5 = 0.5**(A)** 0.5 × 1 = 0.5

(B) 0.166 × 3 = 0.5

(C) 0.25 × 2 = 0.5

(D) 0.6 × 6 = 3.6

16.(A, B, D) For $\begin{matrix} COOH \\ | \\ COOH \end{matrix}$ and $\begin{matrix} COOK \\ | \\ COOH \end{matrix}$

(A) When each behave as reducing agent

$$(COOH)_2 \rightarrow 2CO_2$$

$n\text{-factor} = 2,$

$\text{Equivalent weight} = M / 2 \text{ (as +3 to +4)}$

(B) For neutralization, milliequivalent of one reactant = Milliequivalent of other reactant can be neutralized

$1000 \times 2 = 1000 \times 2$

$\begin{matrix} COOK \\ | \\ COOH \end{matrix}$ is monobasic.

17.(A, B, D) Number of equivalent of H_2SO_4 = Moles × *n*-factor = 1 × 2 = 2

Number of equivalent of $Ca(OH)_2$ = 1 × 2 = 2 (neutralized)

Number of equivalent of NaOH = 2 × 1 = 2 (neutralized)

Number of equivalent of NH_3 = 2 × 1 = 2 (neutralized)

The required reactions are

$$H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$$
$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$
$$2NH_3 + H_2SO_4 \rightarrow 2NH_4^+ + SO_4^{2-}$$

18.(B, C, D)

In the reduction reactions, $H_2C_2O_4$ and $NaHC_2O_4$ are converted to CO_2, so

$$(C^{3+})_2 \rightarrow 2C^{4+} + 2e$$

Therefore, Equivalent weight =Molecular weight /2.

On titration with 1M $KMnO_4$,

$Mn^{7+} + 5e \rightarrow Mn^{2+}$

Therefore,

$M_{equiv} H_2C_2O_4 = M_{equiv} KMnO_4$

$100 \text{ mL} \times 1M \times 2 = V \text{ mL} \times 1 \text{ M} \times 5$

$M_{equiv} NaHC_2O_4 = M_{equiv} KMnO_4$

$100 \text{ mL} \times 1M \times 2 = V \text{ mL} \times 1 \text{ M} \times 5$

So 100 mL of 1M solution of each is neutralized by equal volumes of 1M $KMnO_4$.

<H2>Assertion–Reasoning Type

1.(B) The reaction involved is

$$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl \text{ (phenolphthalein is the indicator)}$$
$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2 \text{ (second end point with } CH_3OH)$$

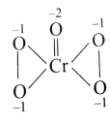
From the above reactions, it is clear that 2 mol of HCl are required for complete neutralization of Na_2CO_3. The titre value with methyl orange corresponds to complete neutralization of Na_2CO_3 and with phenolphthalein correspond to half neutralization of Na_2CO_3.

2.(A) The reaction involved in roasting of FeS_2 is

$$\overset{+2}{Fe}\overset{-2}{S_2} + O_2 \rightarrow \overset{+3}{Fe_2}O_3 + \overset{+4}{S}O_2$$

Since, Fe in +2 state changes to +3 state by loosing one electron. In S_2^{2-} one S is in -2 and another is in 0 state, while changing from $S_2^{2-} \rightarrow \overset{+4}{S}O_2$, both S oxidized to +4 by gaining total 10 electrons. Therefore, the E.W. of FeS_2 is molecular weight/11.

3.(A) The structure of CrO_5 is



Due to the presence of peroxy-linkage in CrO_5, its oxidation number is +6.

4.(D) $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ is a non-redox reaction.

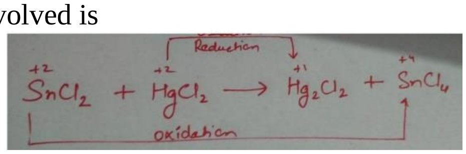
So equivalent weight of $CaCO_3 = \dfrac{\text{Molecular weight}}{2} = 100 = 50$

5.(A) Since, the reduction potential of Fe is more than H_2, therefore, Fe liberated H_2 when reacts with HCl.

6.(D) Hydrogen peroxide acts as an oxidizing as well as reducing agent. Peroxides behave as oxidizing as well reducing agents.

7.(C) Carbon shows +2 oxidation state in HCN.
Here, the contribution of coordinate bond is neglected because the bond is directed from more electronegative nitrogen atom to less electronegative carbon atom.
Carbon shows different oxidation states in different compounds.

8.(A) The reaction involved is



When excess of $SnCl_2$ is added, Hg^{2+} disproportionation easily rather being reduced.

9.(A) Since, moving down a group, the electropositive nature of the elements increase, therefore, I can form interhalogen compounds with F, Cl and Br with positive oxidation states.

10.(C) SO_2 is an effective reducing as well as the oxidizing agent.

$$SO_2 + 2H_2S \rightarrow 3S + 2H_2O$$

<H2>Comprehension Type

1.(C) Since, in the given reaction, Mn simultaneously reduced and oxidized, therefore, this reaction is an example of disproportionation reaction.

$$2\overset{+6}{Mn}O_4^{2-} + 2H_2O \rightarrow \overset{+4}{Mn}O_2 + 2\overset{+7}{Mn}O_4^- + 4OH^-$$

2.(D) $AgCl + NH_3 \rightarrow [Ag(NH_3)_2]Cl$ no change in oxidation state.

3.(B) $\overset{+4}{Mn}O_2 \rightarrow \overset{+6}{M}nO_4^{2-}$ reducing agent.

4.(D) For the given reaction,

$$\overset{+2\ -2}{AX} + \overset{+1\ -1}{BY} + H_2O \rightarrow HA^- + OY^{2+} + \overset{-2}{X_2}\overset{+4}{B}$$

Therefore, both B and Y undergoes oxidation.

5.(D) $\text{Oxidation states } B \le +4; \text{oxidation states of } Y \le 2$

Oxidation states of B = +1, Y = –1

Oxidation states of B = +2, Y = –2

Oxidation states of B = +3, Y = –3

6.(B) 2AX + BY + $2AX + BY + H_2O \rightarrow 2HA + OY + X_2B$

7.(B) Loss of oxygen from OCl^- and loss of hydrogen from $[Cr(OH)_4]^-$.

8.(C) From the given reactions, only reaction given in option **(C)** is balanced reaction.

9.(D) From the given reaction,

$$1e^- + Cu^{2+} \rightarrow Cu^+$$

Therefore, Equivalent weight = $M_B/n = M_B$

10.(D)

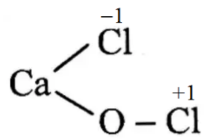
11.(C) $CuSO_4 + KI \rightarrow I_2 \equiv Na_2S_2O_3$

Now, $$\frac{w}{E} \times 1000 = 100 \times 1 \times 1 \Rightarrow w = 15.95$$

So, $\%CuSO_4 = \frac{w}{319} \times 100 = 5\%$

Integer Answer Type

1.(0) From the structure of bleaching powder,



The sum of oxidation states of Cl is 0.

2.(1) $(H_2PO_2)_2^{2-}$

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

3.(7) The balanced redox reaction is

$$3P_4S_6 + 44H^+ + 44NO_3^- \rightarrow 44NO + 12H_3PO_4 + 18SO_2 + 4H_2O$$

4.(3) $MnO_4^- + 2H_2O + 3e^- \quad MnO_2 + 4OH^-$

5.(14) The half reactions are

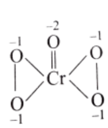
$$2\overset{+5}{N}O_3^- + 8H^+ + 2e^- \rightarrow 2\overset{+3}{N}O^+ + 4H_2O] \times 5$$

$$\overset{+3}{As_2}\overset{-2}{S_3} + 20H_2O \rightarrow 2\overset{+5}{As}O_4^{3-} + 3\overset{+6}{S}O_4^{2-} + 10e^- + 40H^+$$

Overall balanced reaction

$$14\overset{+5}{N}O_3^- + \overset{+3}{As_2}\overset{-2}{S_3} + 16H^+ \rightarrow 14\overset{+3}{N}O^+ + 2\overset{+5}{As}O_4^{3-} + 3\overset{+6}{S}O_4^{2-} + 8H_2O$$

6.(6) The structure of CrO_5 is



Due to the presence of peroxy-linkage in CrO_5, its oxidation number is +6.

7.(2) Milliequivalent of oxalic acid in 16.68 mL = Milliequivalent of NaOH $= 25 \times \frac{1}{15}$

Milliequivalent of oxalic acid in 250 mL $= 25 \times \frac{1}{15} \times \frac{250}{16.68} = 24.98$

Therefore, $\frac{1.575}{(90+18x)/2} \times 1000 = 24.98$

n-Factor of oxalic acid is 2 as dibasic acid. Solving x = 2, that is, $(COOH)_2 \cdot 2H_2O$

<H2>Matrix–Match Type

1.(A) → (r); **(B)** → (s); **(C)** → (p); **(D)** → (q)

(A) $7x + (-18) = 0$ **(B)** $x + 1 - 2 = 0$

(C) $5x - 12 = 0$ **(D)** $+2 + x - 5 + 1 = 0$

2.A → (r, t), B → (p, t), C → (q), D → (s)

$\overset{-3}{N}H_3 \to \overset{+5}{N}O_3^-$ *n*-factor = 8, Equivalent weight = Molecular weight/(+5 – (–3)) = Molecular weight/8

$\overset{+2}{Fe}\overset{+3}{C_2}O_4\overset{+3}{Fe} + 2\overset{+4}{C}O_3^{2-}$ *n*-factor = 3, Equivalent weight = Molecular weight/3

$H_2\overset{+6}{S}O_5 \to \overset{0}{S_8}$ *n*-factor = 6, Equivalent weight = Molecular weight/6

$K\overset{+7}{Mn}O_4 \to \overset{+2}{Mn}$ *n*-factor = 5, Equivalent weight = Molecular weight/5

3.(A) → (s); **(B)** → (r, s); **(C)** → (q, s); **(D)** → (p, s)

(A) $\overset{+5}{V_2}O_5 + 5\overset{0}{Ca} \to 2\overset{0}{V} + 5\overset{+2}{Ca}O$

This is a redox reaction.

(B) $2\overset{+1}{Na}\overset{-1}{H} \to 2\overset{0}{Na} + \overset{0}{H_2}$

This is an example of decomposition as well as redox reaction.

(C) $\overset{0}{P_4} + 3\overset{-}{OH} + 3H_2O \to \overset{-3}{P}H_3 + 3H_2\overset{+2}{P}O_2$

Disproportionation is a reaction in which the same element undergoes oxidation as well as reduction. It is also a type of redox reaction.

(D) $\overset{0}{Ca} + 2\overset{+1}{H_2}\overset{-2}{O} \to \overset{+2}{Ca}(OH)_2 + \overset{0}{H_2}$

This reaction is an example of non-metal displacement reaction. It is also a redox reaction.

4.(A) → (p, r); **(B)** → (r); **(C)** → (p, s); **(D)** → (q, s)

(A) $Zn + 2HCl \to H_2\uparrow + ZnCl_2$

$$2Al + 6HCl \rightarrow 3H_2 \uparrow + 2AlCl_3$$

(B) Aluminium container can be used to store conc. HNO_3 because Al is rendered passive by nitric acid. This is due to oxidation and formation of a thin film of oxide on its surface. Hence Al does not react with conc. HNO_3 so it can be used to store HNO_3.

(C) Zn and Cu can be used as electrode in Daniel cell

$$Zn \rightarrow Zn^{2+} + 2e^- \quad \text{(Oxidation)}$$

$$Cu^{2+} + 2e^- \rightarrow Cu \quad \text{(Reduction)}$$

(D) Silver and copper do not react with dilute acids to give H_2 gas because it acts as oxidizing agent and it has higher reduction potential value.

5.(A) $\rightarrow$ (r); **(B)** $\rightarrow$ (s); **(C)** $\rightarrow$ (q); **(D)** $\rightarrow$ (p, r)

(A) $2x + 2(-2) = 0$ so $x = +2$

(B) $x + (+2) + (-2) \times 4 = 0$ $x = +6$

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

(D) $2x + (-1)2 = 0$ so $x = +1$ and $2(-2) + 2y = 0$ so $y = +2$