DAY NINE

Redox Reactions

Learning & Revision for the Day

- Concepts of Oxidation and Reduction
- Oxidation Number
- Redox Reactions

The transformation of one kind of matter into another takes place through the various types of reactions. Redox reactions are the important category of such reactions. These are the reactions in which oxidation and reduction take place simultaneously.

Concepts of Oxidation and Reduction

Loss of electron by an atom is called **oxidation** or de-electronation, while gain of electron by an atom is called **reduction** or electronation.

- **Oxidants or Oxidising Agents** These are the substances which
  (i) oxidise other, (ii) get reduced,
  (iii) gain electrons (i.e. their oxidation number decreases during a reaction)
  e.g. MgO, CaO, O_2, halogens, H_2SO_4, $KMnO_4$ etc.
- **Reductants or Reducing Agents** These are the substances which
  (i) reduce others, (ii) get oxidised,
  (iii) lose electrons (i.e. their oxidation number increases during a reaction)
  e.g. Na, Al, NaH, LiH, $FeSO_4$, $HgCl_2$ etc.

NOTE The compounds having highest oxidation state in their compounds act as oxidants, while those having lowest oxidation state in their compounds act as reductants.

Equivalent Weights of Oxidising Agent (OA) or Reducing Agent (RA)

It may be formulated as,

$$E_{OA/RA} = \frac{\text{Molar mass of OA / RA agent}}{\text{Number of electrons lost or gained per formula unit of RA / OA}}$$

H_2O_2 is both oxidising and reducing agent but its equivalent weight as either oxidising or reducing agents are the same, i.e. 17.

Oxidation Number

It is defined as the real or imaginary charge, which an atom appears to have in its combined state.

Valency of an element is always a whole number. It can neither be zero nor fractional. While oxidation number may be positive or negative. It can be zero or fractional.

Rules for Assigning Oxidation Number

It can be calculated with the help of following rules:

- The oxidation number of an element in its elementary state is zero, e.g. H in H_2, S in S_8, P in P_4.
- Oxidation number of an ion is equal to the electrical charge present on it.
- Oxidation number of a compound is zero.
- Oxidation number of fluorine is always –1 in all of its compounds.
- The oxidation number of alkali metals is always +1 and those of alkaline earth metals is +2.
- Oxidation number of hydrogen is +1 except in ionic hydrides, where it is –1.
- Two oxidation numbers of N are –3 and +3, when it is bonded with less electronegative and more electronegative atoms respectively.
- Oxidation number of oxygen is –2 except in $OF_2(+2)$, $O_2F_2(+1)$, peroxides (-1) and superoxides $(-1/2)$.
- The oxidation number of halogens is always –1 in metal halides.
- In interhalogen compounds, the more electronegative of the two halogens gets the oxidation number of –1.
- Oxidation number of metals in amalgams and carbonyls, e.g. $[Fe(CO)_5]$ is zero.

Fractional Oxidation States

These are often used to represent the average oxidation states of several atoms in a structure. e.g. In KO_2, the superoxide ion has a charge of –1 divided among two equivalent atoms, so each oxygen is assigned an oxidation state of –1/2. This ion is described as a resonance hybrid of two Lewis structures and each oxygen has oxidation state 0 in one structure and –1 in the other.

- For the cyclopentadienyl ion $C_5H_5^-$, the oxidation state of C is $(-1)+(-1/5)=-6/5$. The –1 occurs because each C is bonded to one less electronegative H and the –1/5 because the total ionic charge is divided among five equivalent C.
- If the average refers to atoms which are not equivalent, the average oxidation state may not be representative of each atom. This is true in magnetite Fe_3O_4, whose formula leads to an average oxidation state of +8/3. Infact one third of the iron ions are Fe^{3+} and two third Fe^{2+}.
- Examples of fractional oxidation states for carbon

  (i) $-(6/7): C_7H_7^+$ (ii) $-(10/8): C_8H_8^{2-}$

Redox Reactions

The reaction, which involves oxidation and reduction as its two half reactions is called redox reaction.

These are of three types as follows:

1. Intermolecular Redox Reactions

In such redox reaction molecule of one reactant is oxidised whereas, molecule of other is reduced.

e.g. $$NO_3^- + H_2S + H_2O + H^+ \longrightarrow NH_4^+ + HSO_4^-$$

Oxidised

Reduced

These are further divided into two types:

- **Combination reactions** are those in which two or molecules (in their zero oxidation state) combine together and one gets oxidised while the other gets reduced.

  e.g. $$\underset{\text{Reductant}}{\overset{0}{C}} + \underset{\text{Oxidant}}{\overset{0}{O_2}} \longrightarrow \overset{+4\,-2}{C\,O_2}$$

- **Displacement reactions** in which an atom or ion in a compound is replaced by an atom or ion.

  These are further divided into two types:

  (i) **Metal displacement reactions** in which metal is displaced.

  $$\underset{\text{Oxidant}}{CuSO_4} + \underset{\text{Reductant}}{Zn} \longrightarrow Cu + ZnSO_4$$

  (ii) **Non-metal displacement reactions** in which non-metal is displaced.

  $$\underset{\text{Reductant}}{2\,Na} + \underset{\text{Oxidant}}{2H_2O} \longrightarrow 2NaOH + H_2$$

2. Intramolecular Redox Reactions

These are the reactions which involve oxidation of one element of a compound as well as reduction of other element of the same compound. Decomposition reactions are also intramolecular redox reactions but to be a redox reaction, it is essential that one of the products of decomposition must be in the elemental state. e.g.

Oxidation

$$(NH_4)_2Cr_2O_7 \xrightarrow{\Delta} N_2 + Cr_2O_3 + 4H_2O$$

Reduction

3. Disproportionation Reactions

These are the reactions which involve oxidation and reduction of the same element, e.g.

$$Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^{-1} + H_2O$$

or

$$\underset{0}{Cl_2} + 2\underset{-2\;+1}{OH^-} \longrightarrow \underset{-1}{Cl^-} + \underset{+1\;-2}{ClO^-} + \underset{+1\;-2}{H_2O}$$

Balancing of Redox Reactions

These reactions can be balanced by two methods:

1. Ion Electron Method

(by loss and gain of electrons)

This method involves the following steps:

- Write redox reaction in ionic form.
- Split redox reaction into oxidation half and reduction half reactions.
- Balance the atoms of each half reaction by using simple multiples.
- For balancing H and O, add H^+ ion and H_2O to the appropriate sides, similarly for basic medium, add OH^- and H_2O to the appropriate sides.
- Balance the charge on both the sides and multiply one or both half reactions by suitable number to equalise number of electrons in both equations.
- Add the two balance half reactions and cancel common terms.

2. Oxidation Number Method

This method involves the following steps:

- Assign oxidation number to the atoms in the equation and write separate equations for atoms undergoing oxidation and reduction.
- Find the change in oxidation number in each equation and make the change equal in both the equations by multiplying with suitable integers. After adding both the equations complete the balancing (by balancing H and O).

DAY PRACTICE SESSION 1

FOUNDATION QUESTIONS EXERCISE

1 Which of the following processes does not involve oxidation of iron? → **CBSE-AIPMT 2015**

(a) Rusting of iron sheets
(b) Decolourisation of blue $CuSO_4$ solution by iron
(c) Formation of $Fe(CO)_5$ from Fe
(d) Liberation of H_2 from steam by iron at high temperature

2 Which of the following does not give oxygen on heating? → **NEET 2013**

(a) $Zn(ClO_3)_2$ (b) $K_2Cr_2O_7$
(c) $(NH_4)_2Cr_2O_7$ (d) $KClO_3$

3 In the reaction,

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

the substance oxidised is

(a) H_2S (b) SO_2 (c) S (d) H_2O

4 A sulphur containing species that cannot be a reducing agent is

(a) SO_2 (b) SO_3^{2-} (c) H_2SO_4 (d) S^{2-}

5 Which of the following halogen acid is better reducing agent?

(a) HCl (b) HBr (c) HI (d) HF

6 Which substance is serving as a reducing agent in the following reaction?

$$14H^+ + Cr_2O_7^{2-} + 3Ni \longrightarrow 2Cr^{3+} + 7H_2O + 3Ni^{2+}$$

(a) H_2O (b) Ni (c) H^+ (d) $Cr_2O_7^{2-}$

7 Which one of the following reactions represents the oxidising property of H_2O_2?

(a) $2KMnO_4 + 3H_2SO_4 + 5H_2O_2 \longrightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 5O_2$
(b) $2K_3[Fe(CN)_6] + 2KOH + H_2O_2 \longrightarrow 2K_4[Fe(CN)_6] + 2H_2O + O_2$
(c) $PbO_2 + H_2O_2 \longrightarrow PbO + H_2O + O_2$
(d) $2KI + H_2SO_4 + H_2O_2 \longrightarrow K_2SO_4 + I_2 + 2H_2O$

8 In the reaction,

$$2KMnO_4 + 16HCl \longrightarrow 5Cl_2 + 2MnCl_2 + 2KCl + 8H_2O$$

the reduction product is

(a) Cl_2 (b) $MnCl_2$
(c) H_2O (d) KCl

9 The reaction in which hydrogen peroxide acts as a reducing agent is

(a) $PbS + 4H_2O_2 \longrightarrow PbSO_4 + 4H_2O$
(b) $2KI + H_2O_2 \longrightarrow 2KOH + I_2$
(c) $Ag_2O + H_2O_2 \longrightarrow 2Ag + H_2O + O_2$
(d) $H_2SO_3 + H_2O_2 \longrightarrow H_2SO_4 + H_2O$

10 In the reaction,

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

(a) bromine is oxidised and carbonate is reduced
(b) bromine is reduced and water is oxidised
(c) bromine is neither reduced nor oxidised
(d) bromine is both reduced and oxidised

11 The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is

(a) one fifth (b) five (c) one (d) two

12 Whenever a reaction between an oxidising agent and a reducing agent is carried out

(a) a compound of lower oxidation state is formed if the reducing agent is in excess

(b) a compound of higher oxidation state is formed if the oxidising agent is in excess
(c) Both (a) and (b)
(d) None of the above

13 $KMnO_4$ can be prepared from K_2MnO_4 as per the reaction

$$3MnO_4^{2-} + 2H_2O \rightleftharpoons 2MnO_4^- + MnO_2 + 4OH^-$$

The reaction can go to completion by removing OH^- ions by adding → **NEET 2013**

(a) KOH (b) CO_2 (c) SO_2 (d) HCl

14 What is the equivalent mass of IO_4^-, when it is converted into I_2 in acid medium?

(a) $\frac{M}{4}$ (b) $\frac{M}{5}$
(c) $\frac{M}{6}$ (d) $\frac{M}{7}$

15 In acidic medium, H_2O_2 changes $Cr_2O_7^{-2}$ to CrO_5 which has two (—O—O—) bonds. Oxidation state of Cr in CrO_5 is → **CBSE-AIPMT 2014**

(a) +5 (b) +3 (c) +6 (d) –10

16 The reaction of aqueous $KMnO_4$ with H_2O_2 in acidic conditions gives → **CBSE-AIPMT 2014**

(a) Mn^{4+} and O_2 (b) Mn^{2+} and O_2
(c) Mn^{2+} and O_3 (d) Mn^{4+} and MnO_2

17 Which is the strongest acid in the following? → **NEET 2013**

(a) $HClO_3$ (b) $HClO_4$ (c) H_2SO_3 (d) H_2SO_4

18 A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction, which element undergoes maximum change in the oxidation number? → **CBSE-AIPMT 2012**

(a) S (b) H (c) Cl (d) C

19 In which of the following compounds, nitrogen exhibits highest oxidation state? → **CBSE-AIPMT 2012**

(a) N_2H_4 (b) NH_3 (c) N_3H (d) NH_2OH

20 When Cl_2 gas reacts with hot and concentrated sodium hydroxide solution, the oxidation number of chlorine changes from → **CBSE-AIPMT 2012**

(a) zero to +1 and zero to –5
(b) zero to –1 and zero to +5
(c) zero to –1 and zero to +3
(d) zero to +1 and zero to –3

21 The oxidation state of chromium in the final product formed by the reaction between KI and acidified potassium dichromate solution is

(a) +3 (b) +2 (c) +6 (d) +4

22 Oxidation states of P in $H_4P_2O_5$, $H_4P_2O_6$, $H_4P_2O_7$, are respectively → **CBSE-AIPMT 2010**

(a) +3, + 5, + 4 (b) +5, + 3, + 4
(c) +5, + 4, + 3 (d) +3, + 4, + 5

23 In which of the following reaction, there is no change in oxidation number

(a) $SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$
(b) $2Na + O_2 \longrightarrow Na_2O_2$
(c) $Na_2O + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$
(d) $4KClO_3 \longrightarrow 3KClO_4 + KCl$

24 Which of the following is not a redox reaction?

(a) $CaCO_3 \longrightarrow CaO + CO_2$
(b) $O_2 + 2H_2 \longrightarrow 2H_2O$
(c) $Na + H_2O \longrightarrow NaOH + \frac{1}{2} H_2$
(d) $MnCl_3 \longrightarrow MnCl_2 + \frac{1}{2} Cl_2$

25 The reaction, $3ClO^-(aq) \longrightarrow ClO_3^-(aq) + 2Cl^-(aq)$ is an example of

(a) oxidation reaction
(b) reduction reaction
(c) disproportionation reaction
(d) decomposition reaction

26 Which is the best description of behaviour of bromine in the reaction given below?

$$H_2O + Br_2 \longrightarrow HBr + HOBr$$

(a) Proton accepted only
(b) Both oxidised and reduced
(c) Oxidised only
(d) Reduced only

27 Which of the following elements does not show disproportionation tendency?

(a) Cl (b) Br
(c) F (d) I

28 The chemical that undergoes self oxidation and self reduction in the same reaction is

(a) benzyl alcohol (b) acetone
(c) formaldehyde (d) acetic acid

29 For the redox reaction,

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

correct stoichiometric coefficient of MnO_4^-, $C_2O_4^{2-}$, H^+ are

(a) 2, 5, 16 (b) 16, 5, 2
(c) 5, 16, 2 (d) 2, 16, 5

30 In the ionic equation,

$$BiO_3^- + 6H^+ + xe^- \longrightarrow Bi^{3+} + 3H_2O$$

The value of x is

(a) 6 (b) 2
(c) 4 (d) 3

31 In the balanced chemical reaction;

$$IO_3^- + aI^- + bH^+ \longrightarrow cH_2O + dI_2$$

a, b, c and d respectively correspond to

(a) 5, 6, 3, 3 (b) 5, 2, 6, 3
(c) 3, 5, 3, 6 (d) 5, 6, 5, 5

32 For the redox reaction,

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

the correct coefficients of the reactants for the balanced equation are → **NEET 2018**

	MnO_4^-	$C_2O_4^{2-}$	H^+
(a)	2	16	5
(b)	2	5	16
(c)	16	5	2
(d)	5	16	2

33 For the redox reaction,

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

In basic medium, coefficients of Zn, NO_3^- and OH^- in the balanced equation respectively, are

(a) 4, 1, 7 (b) 7, 4, 1
(c) 4, 1, 10 (d) 1, 4, 10

34 For decolourisation of 1 mole of $KMnO_4$, the moles of H_2O_2 required is

(a) $\frac{1}{2}$ (b) $\frac{3}{2}$
(c) $\frac{5}{2}$ (d) $\frac{7}{2}$

35 The value of n in the reaction

$$Cr_2O_7^{2-} + 14H^+ + nFe^{2+} \longrightarrow 2Cr^{3+} + nFe^{3+} + 7H_2O$$

will be

(a) 2 (b) 3
(c) 6 (d) 7

DAY PRACTICE SESSION 2

PROGRESSIVE QUESTIONS EXERCISE

1 Thiosulphate reacts differently with iodine and bromine in the reactions given below

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$
$$S_2O_3^{2-} + 2Br_2 + 5H_2O \longrightarrow 2SO_4^{2-} + 4Br^- + 10H^+$$

Which of the following statements justifies the above dual behaviour of thiosulphate?

(a) Bromine is a stronger oxidant than iodine
(b) Bromine is a weaker oxidant than iodine
(c) Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions
(d) Bromine undergoes oxidation and iodine undergoes reduction in these reactions

2 In which of the following, the oxidation number of oxygen has been arranged in the increasing order?

(a) $OF_2 < KO_2 < BaO_2 < O_3$ (b) $BaO_2 < KO_2 < O_3 < OF_2$
(c) $BaO_2 < O_3 < OF_2 < KO_2$ (d) $KO_2 < OF_2 < O_2 < BaO_2$

3 The complex $[Fe(H_2O)_5NO]^{2+}$ is formed in the ring test for nitrate when freshly prepared $FeSO_4$ solution is added to aqueous solution of NO_3^- followed by addition of conc. H_2SO_4. This complex is formed by charge transfer in which

(a) Fe^{2+} changes to Fe^{3+} and NO^+ changes to NO
(b) Fe^{2+} changes to Fe^{3+} and NO changes to NO^+
(c) Fe^{2+} changes to Fe^+ and NO changes to NO^+
(d) no charge transfer takes place

4 While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. This is because

(a) in SO_2 and H_2O_2, S and O are in their lowest oxidation state
(b) in O_3 and HNO_3, N and O are in their highest oxidation state
(c) in O_3 and HNO_3, N and O are in their lowest oxidation state
(d) Both (a) and (b)

5 The brown ring complex compound is formulated as $[Fe(H_2O)_5(NO)]\,SO_4$. The oxidation state of iron is

(a) 1 (b) 2
(c) 3 (d) 0

6 MnO_4^- is a good oxidising agent in different medium changing to

$$\underset{1}{MnO_4^-} \longrightarrow \underset{2}{Mn^{2+}} \longrightarrow \underset{3}{MnO_4^{2-}} \longrightarrow \underset{4}{MnO_2} \longrightarrow \underset{5}{Mn_2O_3}$$

Changes in oxidation number respectively, are

(a) 1, 3, 4, 5 (b) 5, 4, 3, 2
(c) 5, 1, 3, 4 (d) 2, 6, 4, 3

7 Which has least number of equivalent per mole for the reactant?

(a) MnO_4^- changes to MnO_2 (b) MnO_4^- changes to Mn^{2+}
(c) MnO_4^- changes to MnO_4^{2-} (d) MnO_4^- changes to Mn_2O_3

8 The reaction,

$$Cl_2(g) + 2OH^-(aq) \longrightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$$

represents the process of bleaching. Identify the species that bleaches the substances due to its oxidising action.

(a) Cl^- (b) Cl_2
(c) OH^- (d) ClO^-

9 Consider the following redox reactions:

I. $(NH_4)_2Cr_2O_7 \longrightarrow N_2 + Cr_2O_3 + H_2O$

II. $NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

III. $2KClO_3 \longrightarrow 2KCl + 3O_2$

Select the intramolecular redox reactions.

(a) I, II (b) II, III
(c) I, III (d) I, II, III

10 I^- reduces IO_3^- to I_2 and itself oxidised to I_2 in acidic medium. final reaction is

(a) $I^- + IO_3^- + 6H^+ \longrightarrow I_2 + 3H_2O$
(b) $I^- + IO_3^- \longrightarrow I_2 + O_3$
(c) $5I^- + IO_3^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$
(d) None of the above

11 For a titration of 100 cm^3 of 0.1 M Sn^{2+} to Sn^{4+} 50 cm^3 of 0.40 M Ce^{4+} solution was required. The oxidation state of cerium in the reduction product is

(a) +1 (b) +2
(c) +3 (d) 0

12 For the following reaction, consider the following statements:

$$2\,Cr(OH)_3 + 3H_2O_2 + 4OH^- \longrightarrow 2CrO_4^{2-} + 8H_2O$$

I. There is colour change from green precipitate to yellow coloured solution.

II. Oxidation number of Cr changes from + 3 to + 6.

III. Oxidation number of O in H_2O_2 changes from −2 to −1.

Select the correct statements (s).

(a) Only I (b) Both I and II
(c) Only II (d) Both I and III

13 For the redox reaction,

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

In basic medium, coefficients of Zn, NO_3^- and OH^- in the balanced equation are respectively

(a) 7, 4, 1 (b) 4, 1, 10 (c) 1, 4, 10 (d) 4, 1, 7

14 Consider the following reaction,

$$\begin{array}{c} CHO \\ | \\ CHO \end{array} + OH^- \longrightarrow \begin{array}{c} COO^- \\ | \\ CH_2OH \end{array}$$

Select the incorrect statement.

(a) It is a disproportionation reaction
(b) It is intramolecular redox reaction
(c) OH^- is a reducing as well as oxidising agent
(d) $\begin{array}{c} CHO \\ | \\ CHO \end{array}$ is a reducing as well as oxidising agent

15 Select the correct statement about the following reaction,

$$NH_4^+ + NO_2^- \longrightarrow N_2 + 2H_2O$$

(a) oxidation number of N has changed from −2 to + 2
(b) oxidation number of N in NH_4^+ changed from −3 to 0 and that in NO_2^- changed from + 3 to 0.
(c) oxidation number of N in NH_4^+ changed from + 1 to 0 and that in NO_2^- changed from − 1 to 0.
(d) no change in oxidation number

ANSWERS

SESSION 1

1 (c)	**2** (c)	**3** (a)	**4** (c)	**5** (c)	**6** (b)	**7** (d)	**8** (b)	**9** (c)	**10** (d)
11 (d)	**12** (c)	**13** (b)	**14** (d)	**15** (c)	**16** (b)	**17** (b)	**18** (c)	**19** (c)	**20** (b)
21 (a)	**22** (d)	**23** (c)	**24** (a)	**25** (c)	**26** (b)	**27** (c)	**28** (c)	**29** (a)	**30** (b)
31 (a)	**32** (b)	**33** (c)	**34** (c)	**35** (c)					

SESSION 2

1 (a)	**2** (b)	**3** (c)	**4** (b)	**5** (a)	**6** (c)	**7** (c)	**8** (d)	**9** (d)	**10** (c)
11 (c)	**12** (b)	**13** (b)	**14** (c)	**15** (b)					

Hints and Explanations

SESSION 1

1 (a) $\overset{0}{Fe} + H_2O + \underset{\text{From air}}{O_2} \longrightarrow \overset{III}{Fe_2}O_3 \cdot xH_2O$

(b) $\overset{0}{Fe} + CuSO_4 \longrightarrow Fe\overset{II}{S}O_4 + Cu$

(c) $\overset{0}{Fe} + 5CO \longrightarrow \overset{0}{Fe}(CO)_5$

(d) $\overset{0}{Fe} + \underset{\text{Steam}}{H_2O} \longrightarrow \overset{III}{Fe_2}O_3 + H_2$

$\therefore$ Formation of $Fe(CO)_5$ from Fe does not involve oxidation of iron because there is no change in oxidation state.

2

$$Zn(ClO_3)_2 \xrightarrow{\Delta} ZnCl_2 + 3O_2$$

$$2\,K_2Cr_2O_7 \xrightarrow{\Delta} 2K_2CrO_4 + Cr_2O_3 + \frac{3}{2}O_2$$

$$(NH_4)_2Cr_2O_7 \xrightarrow{\Delta} N_2 + Cr_2O_3 + 4H_2O$$

$$2KClO_3 \xrightarrow{\Delta} 2KCl + 3\,O_2$$

3

Reduction

$$\overset{+4}{SO_2} + 2H_2\overset{-2}{S} \longrightarrow 3\,\overset{0}{S} + 2H_2O$$

Oxidation

$\therefore$ H_2S is oxidised in this reaction.

4 If the species is a reducing agent, it means it can be oxidised easily thus, it should have an oxidation number less than maximum values of oxidation number.

S. No.	Species	Oxidation number
(a)	SO_2	4
(b)	SO_3^{2-}	4
(c)	SO_4^{2-}	6
(d)	S^{2-}	–2

SO_2, SO_3^{2-} and S^{2-} can be reducing agents but SO_4^{2-} cannot.

5 The strength of HI bond is least and hence, it acts as a better reducing agent.

6 Ni metal in this reaction acts as good reducing agent.

7 The reaction in which H_2O_2 is reduced, while the other reactant is oxidised, represents the oxidising property of H_2O_2.

Reduction

$$2\overset{-1}{KI} + H_2SO_4 + H_2\overset{-1}{O_2} \longrightarrow K_2SO_4 + \overset{0}{I_2} + 2H_2\overset{-2}{O}$$

Oxidation

8 In this reaction, oxidation number of Mn changes from + 7 in $KMnO_4$ to + 2 in $MnCl_2$. Thus, $MnCl_2$ is the reduction product.

9 In the reaction,

$Ag_2O + H_2O_2 \longrightarrow 2Ag + H_2O + O_2$

Ag_2O is reduced to Ag by H_2O_2.

10 Oxidation number of Br increases from 0 in Br_2 to + 5 in BrO_3^- ion and oxidation number of Br also decreases from 0 (Br_2) to $-1(Br^-)$. Therefore, bromine is oxidised as well as reduced.

11 In alkaline medium, $KMnO_4$ is reduced to MnO_2 (colourless).

$$2KMnO_4 + 2H_2O \longrightarrow 2MnO_2 + 2KOH + 3[O]$$

$$KI + 3[O] \longrightarrow KIO_3,$$

$$2KMnO_4 + H_2O + KI \longrightarrow 2MnO_2 + 2KOH + KIO_3$$

Hence, two moles of $KMnO_4$ are reduced by one mole of KI .

12 Both the statements are correct. This can be proved by following examples: P_4 is a reducing agent and Cl_2 is an oxidising agent.

(a) $\underset{\text{Excess}}{\overset{0}{P_4}(s)} + 6Cl_2(g) \longrightarrow \underset{\text{Higher oxidation state of P}}{4\overset{+3}{P}Cl_3}$

(b) $\overset{0}{P_4}(s) + \underset{\text{Excess}}{10Cl_2(g)} \longrightarrow \underset{\text{Higher oxidation state of P}}{4\overset{+5}{P}Cl_5}$

Therefore, when P_4 (reducing agent) is in excess, PCl_3 is formed in which oxidation state of P is +3 and if Cl_2 (oxidising agent) is in excess, PCl_5 is formed in which oxidation state of P is +5.

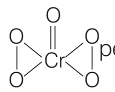
13 HCl and SO_2 are reducing agents, which can reduce MnO_4^-. So, CO_2 which is neither oxidising nor reducing will provide only acidic medium. It can shift reaction in forward direction and reaction can go to completion.

14 $IO_4^- + 7e^- \longrightarrow \frac{1}{2}\overset{0}{I_2}$

Change in oxidation number = 7

Equivalent mass of $IO_4^- = \frac{\text{mol. mass}}{7} = \frac{M}{7}$

15 CrO_5 has butterfly structure having two peroxo bonds are


per oxygen has (–1) oxidation state.

Let, the oxidation state of chromium be 'x'

$\therefore$ $x + 4(-1) + (-2) = 0$, $x = +6$

16 $2KMnO_4 + 5H_2O_2 + 3H_2SO_4 \longrightarrow$

$$K_2SO_4 + 2MnSO_4 + 8H_2O + 5O_2$$

17 In $H\overset{+7}{Cl}O_4$, Cl has highest oxidation number (+7) and its conjugate base is resonance stabilised, hence it is the most acidic.

18 When a mixture of potassium chlorate, oxalic acid and sulphuric acid is heated, the following reaction occurs

$$2K\overset{+5}{Cl}O_3 + H_2\overset{+3}{C_2}O_4 + 2H_2SO_4 \longrightarrow$$

$$\overset{0}{Cl_2} + K_2SO_4 + 10\overset{+4}{C}O_2 + 6H_2O$$

Thus, Cl is the element which undergoes maximum change in the oxidation state.

19 Let the oxidation state of nitrogen in the given compounds be x.

(a) N_2H_4

$$2(x) + (+1)4 = 0, \quad 2x = -4$$
$$\therefore \quad x = -2$$

(b) NH_3

$$x + (+1)3 = 0$$
$$\therefore \quad x = -3$$

(c) N_3H

$$(x)3 + (+1) = 0,$$
$$3x = -1, \quad x = \frac{-1}{3}$$

(d) NH_2OH

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

Thus, oxidation state of nitrogen is highest in N_3H.

20 When chlorine gas reacts with hot and conc. NaOH solution, it disproportionates into chloride (Cl^-) and chlorate (ClO_3^-) ions.

Oxidation

$$\overset{0}{3Cl_2} + \underset{\text{Hot and concentrated}}{6NaOH} \longrightarrow 5Na\overset{-1}{Cl} + Na\overset{+5}{Cl}O_3 + 3H_2O$$

Reduction

In this process, oxidation number of chlorine changes from 0 to -1 and 0 to $+5$.

21 $Cr_2O_7^{2-} + 14H^+ + 6I^- \longrightarrow 2Cr^{3+} + 7H_2O + 3I_2$

$Cr_2O_7^{2-}$ is reduced to Cr^{3+}.

Thus, final oxidation state of Cr is +3.

22 Oxidation state of H is +1 and that of O is -2. Let the oxidation state of P in the given compound is x.

In $H_4P_2O_5$,

$$(+1)\times 4 + 2\times x + (-2)\times 5 = 0$$
$$4 + 2x - 10 = 0$$
$$2x = 6$$
$$\therefore \quad x = +3$$

In $H_4P_2O_6$,

$$(+1)\times 4 + 2\times x + (-2)\times 6 = 0$$
$$4 + 2x - 12 = 0$$
$$2x = 8$$
$$\therefore \quad x = +4$$

In $H_4P_2O_7$,

$$(+1)\times 4 + 2\times x + (-2)\times 7 = 0$$
$$4 + 2x - 14 = 0$$

$\therefore$

$x = +5$

Thus, the oxidation states of P in $H_4P_2O_5$, $H_4P_2O_6$ and $H_4P_2O_7$ are +3, + 4 and +5 respectively.

23 In the reaction

$$\overset{+1}{Na_2}\overset{-2}{O} + \overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4} \longrightarrow \overset{+1}{Na_2}\overset{+6}{S}\overset{-2}{O_4} + \overset{+1}{H_2}\overset{-2}{O}$$

There is no change in odidation number.

24 In redox reaction, oxidation number of elements changes.

But in $\overset{+2}{Ca}CO_3 \longrightarrow \overset{+2}{Ca}O + CO_2$

there is no change in oxidation state. Hence, this is not a redox reaction.

25

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

O.N.increases

O.N.decreases

It is an example of disproportionation reaction.

26 In the reaction,

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

The oxidation number of bromine increases from 0 to + 1 and decreases from 0 to -1, due to this reason, bromine is both oxidised as well as reduced in the above reaction.

27 F is the most electronegative element, so it always shows an oxidation number of -1 and cannot show positive oxidation numbers. In other words, F cannot be simultaneously oxidised as well as reduced, i.e. does not show disproportionation reactions.

28 Formaldehyde is reduced to CH_3OH and it is also oxidised to HCOOH in Cannizzaro reaction in the presence of an alkali (NaOH solution) (i.e. disproportionation reaction).

$$2HCHO + NaOH\ (50\%) \longrightarrow CH_3OH + HCOONa$$

29 $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

30 $BiO_3^- + 6H^+ + xe^- \longrightarrow Bi^{3+} + 3H_2O$

In this reaction, oxidation number of Bi decreases from +5 (BiO_3^-) to + 3 (Bi^{3+}), therefore the value of x =2.

31 Balanced chemical equation is

$$IO_3^- + 5I^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$$

32 The given redox reaction is

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

The reaction can be balanced by cosidering the following steps:

Step I Balance the atoms except H and O.

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

Step II Write the oxidation number of each atom.

Oxidation ($-2e^-$)

$$\overset{+7}{Mn}O_4^- + \overset{+6}{C_2}O_4^{2-} + H^+ \longrightarrow Mn^{2+} + 2\overset{+8}{CO_2} + H_2O$$

Reduction ($+5e^-$)

Step III Cross multiply by change in oxidation number

$$\overset{+7}{Mn}O_4^- \longrightarrow Mn^{2+}\ ;\ 5e^- \text{ gain}$$
$$\overset{+6}{C_2}O_4^{2-} \longrightarrow 2\overset{+8}{CO_2}\ ;\ 2e^- \text{ loss}$$
$$2MnO_4^- + 5C_2O_4^- + H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + H_2O$$

Step IV Balance oxygen by adding H_2O on deficient site.

$$2MnO_4^- + 5C_2O_4^{2-} + H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

Step V Balance hydrogen

$$2MnO_4^{2-} + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

$\therefore$ The coefficients of the reactants, MnO_4^-, $C_2O_4^{2-}$ and H^+ are 2, 5 and 16, respectively.

33 $4Zn + NO_3^- + 7H_2O \longrightarrow 4Zn^{2+} + NH_4^+ + 10OH^-$

34 $\underset{\text{Pink}}{2KMnO_4} + 5H_2O_2 + 3H_2SO_4 \longrightarrow K_2SO_4 + \underset{\text{Colourless}}{2MnSO_4} + 8H_2O + 5O_2$

$\because$ 2 moles of $KMnO_4$ oxidise

$$= 5 \text{ moles of } H_2O_2.$$

$\therefore$ 1 mole of $KMnO_4$ will oxidise

$$= \frac{5}{2} \text{ moles of } H_2O_2$$

35 $Cr_2O_7^{2-} + 14H^+ + nFe^{2+} \longrightarrow 2Cr^{3+} + nFe^{3+} + 7H_2O$

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O \text{ (reduction)} \quad \ldots\text{(i)}$$
$$Fe^{2+} \longrightarrow Fe^{3+} + e^- \text{ (oxidation)} \quad \ldots\text{(ii)}$$

Eq. (ii) is multiplied by 6

$$6Fe^{2+} \longrightarrow 6Fe^{3+} + 6e^-$$

Thus, balanced equation is

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

Hence, the value of n is 6.

SESSION 2

1 $2\overset{+2\,-2}{S_2O_3^{2-}}(aq) + \overset{0}{I_2}(s) \longrightarrow \overset{+2.5}{S_4O_6^{2-}}(aq) + 2I^-(aq)$

$\overset{+2\,-2}{S_2O_3^{2-}}(aq) + 2\overset{0}{Br_2}(l) + 5H_2O(l) \longrightarrow$

$2\overset{+6\,-2}{SO_4^{2-}}(aq) + 4Br^-(aq) + 10H^+(aq)$

Bromine being stronger oxidising agent than I_2, oxidises S of $S_2O_3^{2-}$ to SO_4^{2-} whereas I_2 oxidises it only into $S_4O_6^{2-}$ ion.

2 Let the oxidation number of oxygen in the following compounds be x.

In OF_2, $x + (-1)2 = 0$

$x = +2$

In KO_2, $+1 + (x \times 2) = 0$

$2x = -1, x = -\frac{1}{2}$

In BaO_2, $+2 + (x \times 2) = 0$

$2x = -2, \quad x = -1$

In O_3, oxidation number of oxygen is zero because oxidation number of an element in free state or in any of its allotropic form is always zero.

Thus, the increasing order of oxidation number is

$\underset{-1}{BaO_2} < \underset{-\frac{1}{2}}{KO_2} < \underset{0}{O_3} < \underset{+2}{OF_2}$

3 $Fe^{2+} + e^- \longrightarrow Fe^+$

$NO \longrightarrow NO^+ + e^-$

Fe^+ is formed by the reduction of Fe^{2+} by NO, which is oxidised to NO^+.

4 An element, in its lowest oxidation state, can behave only as reductant, while as oxidant in its highest oxidation state. In intermediate oxidation states, elements can behave as oxidant as well as reductant.

(i) In SO_2, S is in $+4$ oxidation state. It can have minimum oxidation number –2 and maximum oxidation number $+6$. Therefore, S in SO_2 can either decrease or increase its oxidation number. So, SO_2 can act both as oxidising as well as reducing agent.

(ii) In H_2O_2, O is in –1 oxidation state. It can have minimum oxidation number –2 and maximum oxidation number zero ($+1$ and $+2$ also possible in O_2F_2 and OF_2 respectively).

Therefore, 'O' in H_2O_2 can either decrease or increase its oxidation number. So H_2O_2 can act both as oxidising as well as reducing agent.

(iii) In O_3, O is in zero oxidation state. It cannot increase its oxidation number, it can only decrease its oxidation number from zero to –1 or –2. So, ozone can act only as an oxidising agent.

(iv) In HNO_3, the oxidation number of N is $+5$. It is maximum.

So, N in HNO_3 can only decrease its oxidation number. So, it can act as an oxidising agent only.

5 Let oxidation state of Fe in $[Fe(H_2O)_5(NO)]SO_4$ is x

$x + 5(0) + 1 = (-2)$ ($\therefore H_2O$ is a neutral)

$x = +1$

6 $\overset{+7}{MnO_4^-} \longrightarrow Mn^{2+}$ change in oxidation number 5

$\longrightarrow MnO_4^{2-}$ change in oxidation number 1

$\longrightarrow \overset{+4}{MnO_2}$ change in oxidation number 3

$\longrightarrow \overset{+3}{Mn_2O_3}$ change in oxidation number 4

7

	Change in ON	Equivalent per mol of MnO_4^-
(a) $\underset{+7}{MnO_4^-} \longrightarrow \underset{+4}{MnO_2}$	3	3
(b) $\underset{+7}{MnO_4^-} \longrightarrow \underset{+2}{Mn^{2+}}$	5	5
(c) $\underset{+7}{MnO_4^-} \longrightarrow MnO_4^{2-}$	1	1
(c) $\underset{+7}{MnO_4^-} \longrightarrow \frac{1}{2}\underset{3+}{Mn_2O_3}$	4	4

8 $\overset{0}{Cl_2}(g) + 2\overset{-2\,+1}{OH^-}(aq) \longrightarrow \overset{-1\,-2}{ClO^-}(aq) + \overset{-1}{Cl^-}(aq) + \overset{+1\,-2}{H_2O}(l)$

In this reaction, oxidation number of Cl increases from 0 (in Cl_2) to 1 (in ClO^-) as well as decreases from 0 (in Cl_2) to –1 (in Cl^-).

So, it acts as both reducing agent as well as oxidising agent. This is an example of disproportionation reaction.

In this reaction, ClO^- species bleaches the substances due to its oxidising action. [In hypochlorite ion (ClO^-), Cl can decrease its oxidation number from $+1$ to 0 or –1]

9 Intramolecular redox reactions are those, in which different atoms are oxidised and reduced in same substance.

Substance	Oxidised	Reduced
(i) $(NH_4)_2Cr_2O_7$	NH_4^+ to N_2	$Cr_2O_7^{2-}$ to Cr^{3+}
(ii) NH_4NO_3	NH_4^+ to N_2O	NO_3^- to N_2O
(iii) $KClO_3$	O^- to O_2	$ClO_3^-(+5)$ to Cl^-

10 Balanced chemical reaction is

$5I^- + IO_3^- + 6H^+ \rightarrow 3I_2 + 3H_2O$

11 $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$

(increase in oxidation number = 2)

$Ce^{4+} + ne^- \longrightarrow Ce^{(4-n)}$

[decrease in oxidation number = n]

To balance

$nSn^{2+} + 2Ce^{4+} \longrightarrow nSn^{4+} + 2Ce^{(4-n)}$

Millimoles of $Sn^{2+} = 100 \times 0.1 = 10$

Millimoles of $Ce^{4+} = 50 \times 0.4 = 20$

Thus, $\frac{n}{2} = \frac{10}{20}$

$\therefore \quad n = 1$

Thus, oxidation state of cerium in the reduced product

$= Ce^{4-1} = Ce^{3+}$

$= +3$

12 (I) $Cr(OH)_3$ (green precipitate) $\longrightarrow$ CrO_4^{2-} (orange)

(II) $Cr(+3) \longrightarrow Cr(+6)$

(III) Oxidation number of O in H_2O changes from –1 to –2

Thus, I, II are true.

13 Balanced chemical equation is given below:

$4Zn + NO_3^- + 7H_2O \longrightarrow 4Zn^{2+} + NH_4^+ + 10OH^-$

14 One CHO is oxidised to COO^- and one CHO is reduced to CH_2OH.

Thus, it is a disproportionation reaction.

Thus, (a) and (b) is true an

CHO
|
CHO

is reducing as well as oxidising agent.

Thus, (d) is also correct while (c) is incorrect.

15

Oxidation

$\underset{-3}{NH_4^+} + \underset{+3}{NO_2^+} \longrightarrow \underset{0}{N_2} + 2H_2O$

Reduction