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CHEMISTRY REDOX REACTIONS

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REDOX REACTIONS

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REDOX reactions

1. INTRODUCTION

Molecular Equations: $2FeCl_3 + SnCl_2 \rightarrow 2FeCl_2 + SnCl_4$

The reactants and products have been written in molecular forms; thus, the equation is termed as **molecular equation**.

Ionic Equations: The reactions in which the reactants and products are present in the form of ions are called **ionic reactions**.

For example: $2Fe^{3+} + 6Cl^- + Sn^{2+} + 2Cl^- \rightarrow 2Fe^{2+} + 4Cl^- + Sn^{4+} + 4Cl^-$

Or $2Fe^{3+} + Sn^{2+} \rightarrow 2Fe^{2+} + Sn^{4+}$

Illustration 18: Represent the following equation in ionic form. **(JEE MAIN)**

$K_2Cr_2O_7 + 7H_2SO_4 + 6FeSO_4 = 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 7H_2O + K_2SO_4$

Sol: Knowing the oxidation numbers of the elements present, balanced ionic form can be represented. In this equation except H_2O, all are ionic in nature. Representing these compounds in ionic forms,

$$2K^+ + Cr_2O_7{}^{2-} + 14H^+ + 7SO_4^{2-} + 6Fe^{2+} + 6SO_4^{2-} \longrightarrow$$
$$6Fe^{3+} + 9SO_4^{2-} + 2Cr^{3+} + 3SO_4{}^{2-} + 2K^+ + SO_4{}^{2-} + 7H_2O$$

$2K^+$ ions and $13SO_4{}^{2-}$ ions are common on both sides, so these are cancelled. The desired ionic equation reduces to, $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2-} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

Phenomenon of Oxidation and Reduction:

Oxidation or de-electronation is a process which liberates electrons.

Reduction or electronation is a process which gains electrons.

Oxidation	Reduction
a. $M \longrightarrow M^{n+} + ne^-$	$M^{n+} + ne^- \longrightarrow M$
b. $M^{n_1+} \longrightarrow M^{n_2+} + (n_2 - n_1)e^- \quad (n_2 > n_1)$	$M^{n_2+} + (n_2 - n_1)e^- \longrightarrow M^{n_1+} \quad (n_2 > n_1)$
c. $A^{n-} \longrightarrow A + ne^-$	$A + ne^- \longrightarrow A^{n-}$
d. $A^{n_1-} \longrightarrow A^{n_2-} + (n_1 - n_2)e^-$	$A^{n_2^-} + (n_1 - n_2)e^- \longrightarrow A^{n_2^-}$

Note: M may be an atom or a group of atoms; A may be atom or a group of atoms.

Oxidizing and Reducing Agent:

(a) If an element is in its highest possible oxidation state in a compound, it can function as an oxidizing agent, e.g. $KMnO_4$, $K_2Cr_2O_7$, HNO_3, H_2SO_4, $HClO_4$ etc.

(b) If an element is in its lowest possible oxidation state in a compound, it can function as a reducing agent, e.g. H_2S, $FeSO_4$, $Na_2S_2O_3$, $SnCl_2$ etc.

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(c) If an element is in its intermediate oxidation state in a compound, it can function both as an oxidizing agent as well as reducing agent, e.g. H_2O_2, H_2SO_3, HNO_3, SO_2 etc.

(d) If highly electronegative element is in its higher oxidation state in a compound, that compound can function as a powerful oxidizing agent, e.g. $KClO_4$, $KClO_3$, KIO_3 etc.

(e) If an electronegative element is in its lowest possible oxidation state in a compound or in free state, it can function as a powerful reducing agent, e.g. I^-, Br^-, N_3^- etc.

2. MODERN CONCEPT OF OXIDATION AND REDUCTION

According to the modern concept, loss of electrons is oxidation whereas gain of electrons is reduction. Oxidation and reduction can be represented in a general way as shown below:

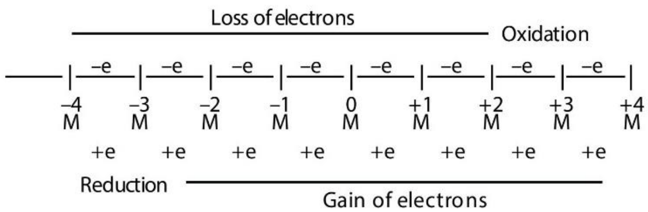


Figure 1.1: Oxidation and Reduction

PLANCESS CONCEPTS

- In a redox process the valency of the involved species changes. The valency of a reducing agent increases while the valency of an oxidising agent decreases in a redox reaction. The valency of a free element is taken as zero.
- Redox reaction involves two half reactions, one involving loss of electron or electrons (oxidation) and the other involving gain of electron or electrons (reduction).

Saurabh Gupta (JEE 2010, AIR 443)

3. ION ELECTRON METHOD FOR BALANCING REDOX REACTIONS

This method involves the following steps:

(a) Divide the complete equations into two half reactions

(i) One representing oxidation

(ii) The other representing reduction

(b) Balance the atoms in each half reaction seperately according to the following steps

(i) Balance all atoms other than oxygen and hydrogen

(ii) To balance oxygen and hydrogen

(c) **Acidic Medium**

(i) Add H_2O to the side which is oxygen deficient to balance oxygen atoms

(ii) Add H^+ to the side which is hydrogen deficient to balance H atoms

REDOX reactions

(d) Basic Medium

(i) Add OH^- to the side which has less negative charge

(ii) Add H_2O to the side which is oxygen deficient to balance oxygen atoms

(iii) Add H^+ to the side which is hydrogen deficient

Illustration 19: $H_2C_2O_4 + KMnO_4 \longrightarrow CO_2 + K_2O + MnO + H_2O$ **(JEE MAIN)**

Sol:

Step 1: Select the oxidant, reductant atoms and write their half reactions, one representing oxidation and other reduction. i.e., $C_2^{+3} \longrightarrow 2C^{+4} + 2e^-$

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

Step 2: Balance the no. of electrons and add the two equation.

$$5C_2^{+3} \longrightarrow 10C^{+4} + 10e^-$$

$$\frac{10e^- + 2Mn^{+7} \longrightarrow 2Mn^{+3}}{5C_2^{+3} + 2Mn^{+7} \longrightarrow 10C^{+4} + 2Mn^{+2}}$$

Step 3: Write complete molecule of the reductant and oxidant from which respective redox atoms were obtained.

$$5H_2C_2O_4 + 2KMnO_4 \longrightarrow 10CO_2 + 2MnO$$

Step 4: Balance other atoms if any (except H and O).

In above example K is unbalanced, therefore,

$$5H_2C_2O_4 + 2KMnO_4 \longrightarrow 10CO_2 + 2MnO + K_2O \text{ (Mentioned as product)}$$

Step 5: Balance O atom using H_2O on desired side.

$$5H_2C_2O_4 + 2KMnO_4 \longrightarrow 10CO_2 + 2MnO + H_2O + 5H_2O$$

4. OXIDATION STATE AND OXIDATION NUMBER

4.1 Oxidation State

It is defined as the charge (real or imaginary) which an atom appears to have when it is in combination. In the case of electrovalent compounds, the oxidation number of an element or radical is the same as the charge on the ion.

4.2 Oxidation Number

(a) Oxidation number of an element in a particular compound represents the number of electrons lost or gained by an element during its change from free state into that compound or Oxidation number of an element in a particular compound represent the extent of oxidation or reduction of an element during its change from free state into that compound.

(b) Oxidation number is given positive sign if electrons are lost. Oxidation number is given negative sign if electrons are gained.

(c) Oxidation number represent real change in case of ionic compounds. However, in covalent compounds it represents imaginary charge.

REDOX reactions

Rules for Calculation of Oxidation Number:

Following rules have been arbitrarily adopted to decide oxidation number of elements on the basis of their periodic properties.

(a) In uncombined state or free state, oxidation number of an element is zero.

(b) In combined state oxidation number of-

(i) F is always –1.

(ii) O is –2. In peroxide it is –1, in superoxides it is –1/2. However in F_2O it is +2.

(iii) H is +1. In ionic hydrides it is –1. (i.e., IA, IIA and IIIA metals).

(iv) Halogens as halide is always –1.

(v) Sulphur as sulphide is always –2.

(vi) Metal is always +ve.

(vii) Alkali metals (i.e., IA group – Li, Na, K, Rb, Cs, Fr) is always +1.

(viii) Alkaline earth metals (i.e., IIA group – Be, Mg, Ca, Sr, Ba, Ra) is always +2.

(c) The algebraic sum of the oxidation number of all the atoms in a compound is equal to zero. e.g. $KMnO_4$.

Ox. no. of K + Ox. no. of Mn + (Ox. no. of O) × 4 = 0

(+1) + (+7) + 4x (–2) = 0

(d) The algebraic sum of all the oxidation no. of elements in a radical is equal to the net charge on the radical. e.g. CO_3^{-2}.

Oxidation no. of C + 3 × (Oxidation no. of O) = –2(4) + 3x (–2) = –2

(e) Oxidation number can be zero, +ve, –ve (integer or fraction)

(f) Maximum oxidation no. of an element is = Group no. (Except O and F)

Minimum oxidation no. of an element is = Group no. –8 (Except metals)

Redox reactions involve oxidation and reduction both. Oxidation means loss of electrons and reduction means gain of electrons. Thus redox reactions involve electron transfer and the number of electrons lost are same as the number of electrons gained during the reaction. This aspect of redox reaction can serve as the basis of a pattern for balancing redox reactions.

Oxidation number of Mn in $KMnO_4$: Let the oxidation number of Mn be x. Now we know that the oxidation numbers of K is +1 and that of O is –2.

$$\begin{array}{ccc} K & Mn & O_4 \\ +1 & +x & +4\times-2 \end{array} \quad \text{or} \quad \begin{array}{ccc} K & Mn & O_4 \\ +1 & +x & -8 \end{array}$$

Now to the sum of oxidation numbers of all atoms in the formula of the compound must be zero, i.e. +1 +x – 8 = 0. Hence, the oxidation number of Mn in $KMnO_4$ is +7.

Illustration 20: What is the oxidation number of Cr in $K_2Cr_2O_7$? **(JEE MAIN)**

Sol: Let the Ox. no. of Cr in $K_2Cr_2O_7$ be x.

We know that, Ox. no. of K = +1

Ox. no. of O = –2

So, 2(Ox. no. K) + 2(Ox. no. Cr) + 7(Ox. no. O) = 0

2(+1) 2(x) 7(–2) = 0

or +2 + 2x – 14 = 0

or $2x = +14 - 2 = +12$

or $x = +\frac{12}{2} = +6$ Hence, oxidation number of Cr in is +6.

Illustration 21: H_2S act only as reductant, whereas SO_2 acts as oxidant and reductant both. **(JEE ADVANCED)**

Sol: Oxidation number of S is –2 in H_2S. It can increase only oxidation number up to +6.

Oxidation number of S is +4 in SO_2. It can increase or decrease as it lies between maximum (+6) and minimum (–2) oxidation number of S.

Illustration 22: Which compound amongst the following has the highest oxidation number of Mn?
$KMnO_4$, K_2MnO_4, MnO_2 and Mn_2O_3. **(JEE MAIN)**

Sol:

		Ox. no. of Mn
$KMnO_4$	+1+x-8=0 x=+7	+7
K_2MnO_4	+2+x-8=0 x=+6	+6
MnO_2	x-4=0 x=+4	+4
Mn_2O_3	2x-6=0 x=+3	+3

Thus, the highest oxidation number for Mn is in $KMnO_4$.

4.3 Balancing of Redox Reactions by Oxidation State Method

This method is based on the fact that the number of electrons gained during reduction must be equal to the number of electrons lost during oxidation. Following steps must be followed while balancing redox equations by this method.

(a) Write the skeleton equation (if not given, frame it) representing the chemical change.

(b) With the help of oxidation number of elements, find out which atom is undergoing oxidation/reduction, and white separate equations for the atom undergoing oxidation/reduction.

(c) Add the respective electrons on the right for oxidation and on the left for reduction equation. Note that the net charge on the left and right side should be equal.

(d) Multiply the oxidation and reduction reactions by suitable integers so that total electrons lost in one reaction is equal to the total electrons gained by other reaction.

(e) Transfer the coefficients of the oxidizing and reducing agents and their products as determined in the above step to the concerned molecule or ion.

(f) By inspection, supply the proper coefficient for the other formulae of substances not undergoing oxidation and reduction to balance the equation.

Illustration 23: $Cr_2O_7^{2-} + I^- + H^+ \longrightarrow Cr^{3+} + I_2$ **(JEE MAIN)**

Sol: (i) Find the oxidation state of atoms undergoing redox change

$$\overset{+6\times 2}{Cr_2}O_7^{2-} + \overset{-1}{I^-} \longrightarrow \overset{+3}{Cr^{3+}} + \overset{0}{I_2}$$

REDOX reactions

(ii) Balance the number of atoms undergoing redox change.

$$\overset{(+6)\times 2}{Cr_2}\,O_7^{2-} + \overset{2\times(-1)}{2I^-} \longrightarrow \overset{(+3)\times 2}{2Cr^{3+}} + \overset{0\times 2}{I_2}$$

(iii) Find the change in oxidation state and balance the change in oxidation states by multiplying the species with a suitable integer.

$$\overset{+12}{Cr_2}\,O_7^{2-} + \overset{-2}{2I^-} \longrightarrow \overset{+6}{2Cr^{3+}} + \overset{0}{I_2}$$

Change in ox. state $= 6$ | Change in ox. state $= 2 \times 3$

As the decrease in oxidation state if chromium is 6 and increase in oxidation state of iodine is 2, so we will have to multiply I^- / I_2 by 3 equalize the changes in oxidation state.

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

(iv) Find the total charges on both the sides and also find the difference of charges.

Charge on LHS $= -2 + 6 \times (-1) = -8$

Charge on RHS $= 2 \times (+3) = +6$

Difference in charge $= +6 - (-8) = 14$

(v) Now, as the reaction is taking place in acidic medium, we will have to add the ions, to H^+ the side falling short in positive charges, so we will add $14H^+$ and LHs to equalize the charges on both sides.

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

(vi) To equalize the H and O atoms, add $7H_2O$ on RHS

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

Illustration 24: Balance the following equation by oxidation number method:

$$Cl_2 + IO_3^- + OH^- \longrightarrow IO_4^- + Cl^- + H_2O$$ **(JEE ADVANCED)**

Sol: Writing oxidation numbers of all atoms,

$$\overset{0}{Cl_2} + \overset{+5}{I}\,\overset{-2}{O_3^-} + \overset{-2}{O}\,\overset{+1}{H^-} \longrightarrow \overset{+7}{I}\,\overset{-2}{O_4^-} + \overset{-1}{Cl^-} + \overset{+1}{H_2}\,\overset{-2}{O}$$

Oxidation numbers of Cl and I have changed.

$$\overset{0}{Cl_2} \longrightarrow 2\overset{-1}{Cl^-} \quad ...(i)$$

$$\overset{+5}{I}O_3^- \longrightarrow \overset{+7}{I}O_4^- \quad ...(ii)$$

Decrease in Ox. no. of Cl = 2 units per Cl_2 molecule

Increase in Ox. no. of I = 2 units per IO_3^- molecule

$$Cl_2 + IO_3^- \longrightarrow IO_4^- + 2Cl^-$$

To balance oxygen, $2OH^-$ ions be added on LHS and one H_2O molecule on RHS. Hence, the balanced equation is

$$Cl_2 + IO_3^- + 2OH^- \longrightarrow IO_4^- + 2Cl^- + H_2O$$

REDOX reactions

5. TYPES OF REACTIONS

The redox reactions are of the following types:

(a) Combination reactions: A compound is formed by chemical combination of two or more elements. The combination of an element or compound with oxygen is called combustion. The combustion and several other combinations which involve change in oxidation state are called redox reactions.

e.g.,

$$\overset{-4}{C}\overset{+1}{H_4} + 2\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\overset{+1}{H_2}\overset{-2}{O}$$

$$\overset{0}{C}(s) + \overset{0}{O_2}(g) \longrightarrow \overset{+4}{C}\overset{-2}{O_2}(g)$$

$$3\overset{0}{Mg} + \overset{0}{N_2} \longrightarrow \overset{+2}{Mg_3}\overset{-3}{N_2}$$

$$\overset{0}{H_2} + \overset{0}{Cl_2} \longrightarrow 2\overset{+1}{H}\overset{-1}{Cl}$$

(b) Decomposition reactions: Decomposition is the reverse process of combination, it involves the breakdown of the compound into two or more components. The product of decomposition must contain at least one component in elemental state.

e.g.,

$$2\overset{+1}{H_2}\overset{-2}{O}(g) \xrightarrow{\Delta} 2\overset{0}{H_2}(g) + \overset{0}{O_2}(g)\,; \qquad 2NaH(s) \xrightarrow{\Delta} 2\overset{0}{Na}(s) + \overset{0}{H_2}(g)$$

$$2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3}(s) \longrightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}(g)$$

In above example, there is no change in oxidation state of potassium. Thus, it should be noted that the decomposition does not result into change in the oxidation number of each element.

(c) Displacement reactions: The reactions in which an atom or ion in a compound is displaced by another atom or ion are called displacement reactions. The displacement reactions are of 2 types:

(i) Metal displacement: In these reactions, a metal in a compound is replace by another metal in an uncombined state. It is found that a metal with stronger reducing character can displace the other metal having a weaker reducing character.

e.g.,

$$\overset{+3}{Cr_2}\overset{-2}{O_3} + 2\overset{0}{Al}(s) \longrightarrow \overset{+3}{Al_2}\overset{-2}{O_3}(s) + 2\overset{0}{Cr}(s)$$

$$\overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O_4} + \overset{0}{Zn}(s) \longrightarrow \overset{+2}{Zn}\overset{+6}{S}\overset{-2}{O_4}(aq) + \overset{0}{Cu}(s)$$

(ii) Non-metal displacement: These displacement reactions generally involve redox reactions, where the hydrogen is displaced. Alkali and alkaline earth metals are highly electropositive, they displace hydrogen from cold water.

$$2\overset{0}{Na}(s) + 2\overset{+1}{H_2}\overset{-2}{O}(l) \longrightarrow 2\overset{+1}{Na}\overset{-2}{O}\overset{+1}{H}(aq) + \overset{0}{H_2}(g)$$

$$\overset{0}{Ca}(s) + 2\overset{+1}{H_2}\overset{-2}{O}(l) \longrightarrow \overset{+1}{Ca}(\overset{-2}{O}\overset{+1}{H})_2(aq) + \overset{0}{H_2}(g)$$

(d) Disproportionation and Oxidation–Reduction: One and the same substance may act simultaneously as an oxidizing agent with the result that a part of it gets oxidized to a higher state and rest of it is reduced to lower state of oxidation. Such a reaction, in which a substance undergoes simultaneous oxidation and reduction is called disproportionation and the substance is said to **disproportionate**.

REDOX reactions

The following are some of the examples of disproportionation:

(a)

Increase

$$\overset{-1}{H_2O_2} + H_2\overset{-1}{O_2} \longrightarrow 2H_2\overset{-2}{O} + \overset{0}{O_2}$$

Decrease

(b)

Decrease

$$3\overset{0}{Cl_2} + \underset{\text{(conc.)}}{6NaOH} \xrightarrow{\text{Hot}} 5Na\overset{-1}{Cl} + Na\overset{+5}{Cl}O_3 + 3H_2O$$

Increase

(e) Oxidation state of chlorine lies between –1 to +7; thus out of ClO^-, ClO_2^-, ClO_3, ClO_4^-; ClO_4^- does not undergo disproportionation because in this oxidation state of chlorine is highest, i.e., +7. Disproportionation of the other oxoanions are:

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

$$6\overset{+3}{Cl}O_2^- \longrightarrow 4\overset{+5}{Cl}O_3^- + 2\overset{-1}{Cl^-} \; ; \qquad 4\overset{+5}{Cl}O_3^- \longrightarrow \overset{-1}{Cl^-} + 3\overset{+7}{Cl}O_4^-$$

Redox reactions

FORMULAE SHEET

Mass

- **Atomic mass**
  Atomic mass $= \dfrac{\text{Average mass of an atom}}{1/12 \times \text{Mass of an atom of } C^{12}}$
- **Molecular mass**
  Molecular mass = The sum of the atomic masses of each atom in the molecule
- **Equivalent weight**
  Equivalent weight $= \dfrac{\text{Molecular weight}}{\text{Valency factor or n factor}}$

Methods

Atomic mass

(i) Dulong and Petit's method
At mass x Sp.heat (cal/gram) = 6.4 (approx.)

(ii) Vapour density method
-suitable for elements whose chlorides are volatile.

(iii) Valency of the element
$= \dfrac{\text{Molecular mass of chloride}}{\text{Equivalent mass metal} + 35.5}$
$= \dfrac{2 \times \text{vapour density of chloride}}{\text{Equivalent mass metal} + 35.5}$

Atomic mass = Equivalent mass of metal x Valency

Molecular mass

(i) Dulong Method
(For Gases)- The ratio of rates of diffusion of two gases is inversely proportional to the square root of their molecular mass or densities.

$$\frac{r_1}{r_1} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}}$$

(ii) Vapour density method (For gases only)-
Mass of a fixed volume of the vapour is compared with the mass of the same volume of hydrogen under same condition; The ratio of these masses is called vapour density or Relative desity.

(iii) Molecular mass = 2 x Vapour density

Equivalent weight

(i) EM of an element
$= \dfrac{\text{Atomic mass}}{\text{Valency}}$

(ii) EM of an acid
$= \dfrac{\text{Molecular mass}}{\text{Basicity}}$
(Basicity of acid is the number of replaceable hydrogen atom in one molecule of the acid).

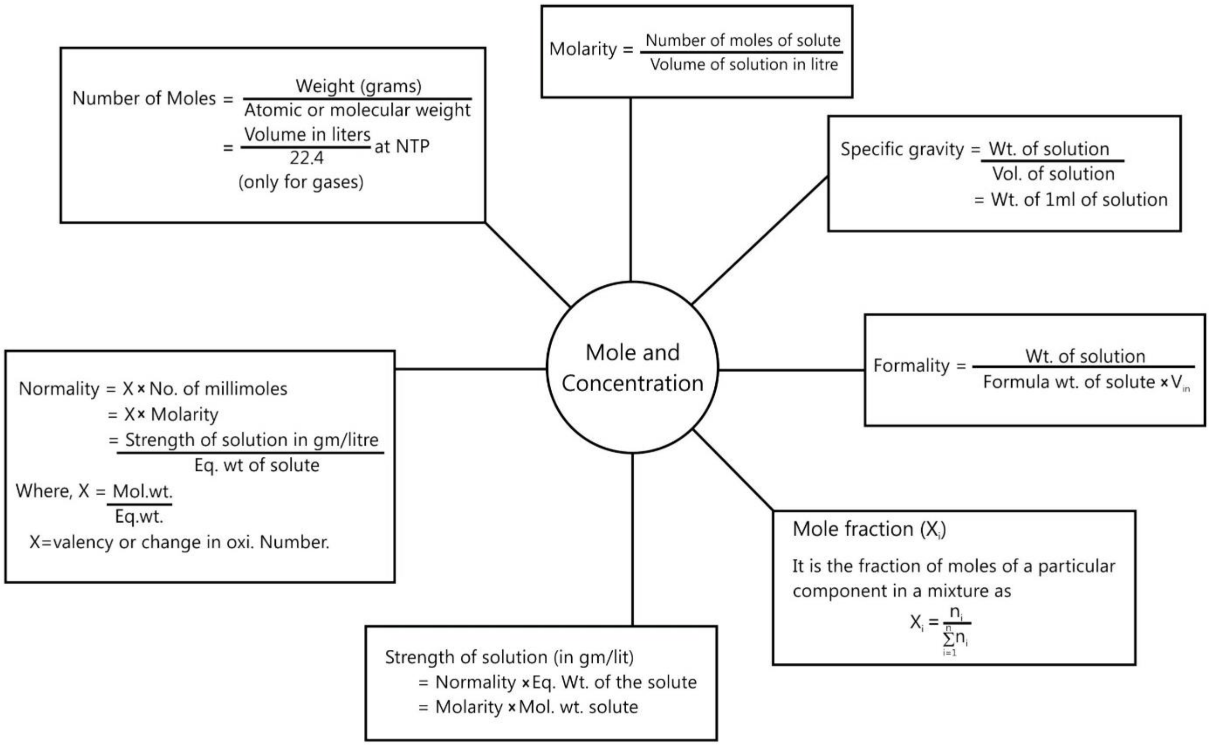
$= \dfrac{\text{Molecular mass}}{\text{Acidity}}$
(Acidity of a base is the number of replaceable OH group in one molecule of the base).

(iv) EM of an salt = $\dfrac{\text{Formula mass}}{\text{Total positive or negative charge}}$

(v) EM of an oxidizing agent=
$\dfrac{\text{Formula mass}}{\text{Number of electrons gained per molecular or Total O.N.}}$

Equivalent mass of common oxidizing agent changes with the medium of the reaction.

REDOX reactions



RULES IN BRIEF

The following are the definitions of 'mole' represented in the form of equations:

(a) Number of moles of molecules = $\frac{\text{Weight in g}}{\text{Molecular weight}}$

(b) Number of moles of atoms = $\frac{\text{Weight in g}}{\text{Atomic weight}}$

(c) Number of moles of gases = $\frac{\text{Volume at NTP}}{\text{Standard molar volume}}$

(Standard molar volume is the volume occupied by 1 mole of any gas at NTP, which is equal to 22.4 litres.)

(d) Number of moles of atoms / molecules / ions / electrons = $\frac{\text{No. of atoms / molecules / ions / electrons}}{\text{Avogadro constant}}$

(e) Number of moles of solute = Molarity × Volume of solution in litres

Or No. of millimoles = Molarity × Volume in mL.

$$\frac{\text{Millimoles}}{1000} = \text{moles}$$

(f) For a compound M_x, N_y, x moles of N = y moles of M

Redox reactions

Reactions

- **Limiting Reactant**

  It is the reactant that is consumed completely during a chemical reaction. If the supplied mass ratio of reactant is consumed completely leaving parts of other uncreated, one that is consumed completely is known as ''limiting reactant''. It determines the amount of product in a given chemical reaction.

- **Mixing of two or more solutions of different molarities**

  If two or more solution of different molarities M_1, M_2, M_3,..... are mixed together, molarity of the resulting solution can be worked out as

  $$M = \frac{M_1V_1+M_2V_2+M_3V_3....}{V_1+V_2+V_3....}$$

REDOX reactions

Quick Revision

The reactions that involve oxidation and reduction as its two half reactions are called **redox reactions**.

1. **Classical Idea of Redox Reactions**
   According to classical concepts, oxidation and reduction are defined as the process that involve:
   Oxidation
   (i) Addition of oxygen
   (ii) Addition of electronegative element
   (iii) Removal of hydrogen
   (iv) Removal of electropositive element
   Reduction
   (i) Removal of oxygen
   (ii) Removal of electronegative element
   (iii) Addition of hydrogen
   (iv) Addition of electropositive element
2. **Oxidising and Reducing Agents**
   In a redox reaction, the substance which oxidises the other species or itself undergo reduction is called the oxidising agent. The substance that reduces the other species and itself undergo oxidation is called the reducing agent.
3. **Redox Reactions in Term of Electron Transfer**
   Loss of electrons or an increase in oxidation state is called oxidation. Gain of electrons or a decrease in oxidation state is called reduction. Because of simultaneous loss and gain of electrons in oxidation-reduction processes, the redox reactions, (or the oxidation-reduction reactions) are also called electron transfer reactions.
4. **Oxidation Number**
   It is defined as "the charge that an atom of the element possesses in its ion or appear to have when present in the combined state with other atoms."
5. **Rules for Calculating Oxidation Number**
   These rules are given below:

***Rule* 1** The oxidation number of an atom in its free or elementary state or in any of its allotropes is zero. e.g. The oxidation state of H in H_2, S in S_8, P in P_4.

***Rule* 2** In case of ions having only one kind of atoms, the oxidation number of each atom is equal to charge present on the ion. e.g. In case of Na^+, Mg^{2+}, Fe^{3+}, Cl^- and O^{2-}, the oxidation state is respectively $+1, +2, +3, -1, -2$.

***Rule* 3** The oxidation state of alkali metals in all their compounds is always + 1. Similarly, in case of alkaline earth metals, it is always +2. For aluminium, oxidation state is always + 3.

***Rule* 4** The oxidation state of oxygen in most of its compounds is – 2, with an exception of peroxides and superoxides in which the oxidation state of oxygen is respectively – 1 and – 1/2.

***Rule* 5** The oxidation state of hydrogen is generally +1 with an exception of metallic hydrides like NaH, CaH_2 etc. In these hydrides, oxidation state of hydrogen is –1 .

***Rule* 6** The oxidation state of fluorine in all of its compounds is always – 1. Other halogens (i.e. chlorine, bromine and iodine) also exhibit – 1 oxidation state but it is not always true.

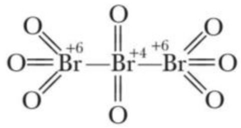
In case of oxoacids and oxoanions, halogens (except fluorine) exhibit positive oxidation state.

***Rule* 7** The algebraic sum of the oxidation numbers of all the atoms present in a compound must be equal to zero.

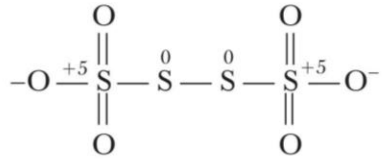
***Rule* 8** In case of polyatomic ions, the algebraic sum of oxidation number of all the atoms present in the ion must be equal to the charge on the ion. e.g. In case of carbonate ion (CO_3^{2-}), it is equal to –2.

6. **Paradox of Fractional Oxidation Number**

- In C_3O_2, $\overset{-2}{O}=\overset{+2}{C}=\overset{0}{C}=\overset{+2}{C}=\overset{-2}{O}$
  two carbon atoms are present in +2 oxidation state each whereas third one is present in zero oxidation state.
  The average of oxidation states of 3 C-atoms in
  $$C_3O_2 = \frac{2+2+0}{3} = \frac{4}{3}$$
- Average oxidation state of Br in Br_3O_8 is 16/3 while the oxidation states of three Br atoms are + 6, + 4 and + 6 as shown below:



- The average oxidation state of four S-atoms in $S_4O_6^{2-}$ is 2.5 while the actual oxidation state of the four S-atoms are +5, 0, 0 and +5 as shown below:



7. **Balancing of Redox Reactions**

(i) **Ion electron method** The method involves the following steps:

(a) Write redox reaction in ionic form.

(b) Split redox reaction into oxidation half and reduction half reactions. Balance atoms of each half reactions by using simple multiples.

For balancing H and O, add H^+ ion and H_2O to the appropriate sides, similarly add OH^- and H_2O to the appropriate sides.

Balance the charge on both 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.

(ii) **Oxidation number method** The method involves the following steps:

(a) Assign oxidation number to the atoms in the equation and write separate equations for atoms undergoing oxidation and reduction.

(b) 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).

REDOX reactions

Objective Questions

Multiple Choose Questions

1. Which of the following reactions represent(s) redox process?

(a) Electrochemical process for extraction of highly reactive metals and non-metals
(b) Manufacturing of caustic soda
(c) Corrosion of metals
(d) All of the above

2. Which of the following processes take place in oxidation?

(a) Addition of oxygen (b) Addition of hydrogen
(c) Removal of oxygen (d) Removal of chlorine

3. Which of the following processes take place in reduction?

(a) Removal of oxygen
(b) Addition of hydrogen
(c) Removal of hydrogen
(d) Both (a) and (b)

4. In the given reaction,

$$2FeCl_3(aq) + H_2(g) \longrightarrow 2FeCl_2(aq) + 2HCl(aq)$$

ferric chloride undergoes

(a) reduction process
(b) oxidation process
(c) addition process
(d) All of the above

5. In the reaction given below, identify the species undergoing oxidation and reduction, respectively

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(a) H_2S is oxidised and Cl_2 is reduced
(b) H_2S is reduced and Cl_2 is oxidised
(c) Both H_2S and Cl_2 are oxidised
(d) Both H_2S and Cl_2 are reduced

6. In the given reaction,

$$2K_4[Fe(CN)_6](aq) + H_2O_2(aq) \longrightarrow 2K_3[Fe(CN)_6](aq) + 2KOH(aq)$$

Which of the following processes takes place?

(a) Oxidation due to removal of potassium
(b) Oxidation due to removal of iron
(c) Reduction due to removal of potassium
(d) Oxidation due to removal of electronegative element

7. In the following reaction,

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

the process taking place w.r.t. Mg is known as

(a) oxidation (b) reduction
(c) redox reaction (d) None of these

8. In the given reaction,

$$CH_2{=}CH_2 + H_2 \longrightarrow H_3C{-}CH_3$$

there occurs

(a) oxidation of ethylene
(b) reduction of ethylene
(c) Both (a) and (b)
(d) None of the above

9. In oxidation process,

(a) oxidation number decreases
(b) number of electrons decreases
(c) oxygen content decreases
(d) number of ions decreases

10. Given the reaction for the discharge of a cobalt-cadmium battery

$$2Co(OH)_3 + Cd + 2H_2O \longrightarrow 2Co(OH)_2 + Cd(OH)_2$$

Which species is oxidised during the discharge of the battery?

(a) Co^{3+} (b) Co^{2+} (c) Cd (d) Cd^{2+}

11. Both oxidation and reduction takes place in

(a) $NaBr + HCl \longrightarrow NaCl + HBr$
(b) $HBr + AgNO_3 \longrightarrow AgBr + HNO_3$
(c) $H_2 + Br_2 \longrightarrow 2HBr$
(d) $CaO + H_2SO_4 \longrightarrow CaSO_4 + H_2O$

Redox reactions

12. In the reaction,

$$2KClO_3 \longrightarrow 2KCl + 3O_2,$$

the elements which have been oxidised and reduced respectively are

(a) chlorine and oxygen
(b) oxygen and chlorine
(c) potassium and oxygen
(d) oxygen and potassium

13. The compound that can work both as an oxidising and reducing agent is

(a) $KMnO_4$ (b) H_2O_2
(c) $Fe_2(SO_4)_3$ (d) $K_2Cr_2O_7$

14. Water molecule is formed by the reaction,

$$2H_2 + O_2 \longrightarrow 2H_2O$$

What does happen in this reaction?

(a) Electrons are transferred from H to O-atom
(b) Electrons are transferred from O to H-atom
(c) Electrons are accepted by H from O-atom
(d) Electrons are donated by O to H-atom

15. Which of the following is not an example of redox reaction? *(NCERT Exemplar)*

(a) $CuO + H_2 \longrightarrow Cu + H_2O$
(b) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
(c) $2K + F_2 \longrightarrow 2KF$
(d) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

16. Match the Column I (Reaction with underlined species) with Column II (Type of change shown by underlined species) and choose the correct option from the codes given below.

Column I (Reactions)	Column II (Type of change)
A. $\underline{2Mg} + O_2 \longrightarrow 2MgO$	1. Removal of hydrogen.
B. $\underline{Mg} + Cl_2 \longrightarrow MgCl_2$	2. Removal of electropositive element.
C. $\underline{2H_2S} + O_2 \longrightarrow 2S + 2H_2O$	3. Addition of oxygen.
D. $\underline{2KI} + H_2O + O_3 \longrightarrow 2KOH + I_2 + O_2$	4. Addition of electronegative element.

Codes

	A	B	C	D
(a)	2	3	4	1
(b)	3	4	1	2
(c)	3	4	2	1
(d)	3	2	1	4

17. Strongest reducing agent is

(a) K (b) Mg (c) Al (d) Ba

18. In the reaction,

$$4Na + O_2 \longrightarrow 2Na_2O,$$

sodium acts as a/an

(a) oxidising agent (b) reducing agent
(c) complexing agent (d) None of these

19. Which reaction indicates the action of HNO_3 as oxidising agent?

(a) $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$
(b) $Ca(OH)_2 + 2HNO_3 \longrightarrow Ca(NO_3)_2 + 2H_2O$
(c) $C_6H_6 + HNO_3 \longrightarrow C_6H_5NO_2 + H_2O$
(d) $NaCl + HNO_3 \longrightarrow HCl + NaNO_3$

20. In the following reaction reducing agent is

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

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

21. Which is the best description behaviour of bromine in the given equation?

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

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

22. In which of the following compounds, nitrogen exhibits highest oxidation state?

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

23. Oxidation states of X, Y, Z are +2, +5 and –2 respectively. The formula of the compound formed by these will be

(a) X_2YZ_6 (b) XY_2Z_6
(c) XY_5 (d) X_3YZ_4

REDOX reactions

24. In which of the following, Fe exhibits minimum oxidation state?

(a) $K_4Fe(CN)_6$
(b) Fe_3O_4
(c) $Fe(CO)_5$
(d) $FeSO_4(NH_4)_2SO_4.6H_2O$

25. Which of the following have been arranged in the decreasing order of oxidation number of sulphur?

(a) $Na_2S_4O_6 > H_2S_2O_7 > Na_2S_2O_3 > S_8$
(b) $H_2SO_4 > SO_2 > H_2S > H_2S_2O_8$
(c) $SO_2^{2-} > SO_4^{2-} > SO_3^{2-} > HSO_4^-$
(d) $H_2SO_5 > H_2SO_3 > SCl_2 > H_2S$

26. In which of the following compounds, an element exhibits two different oxidation states?

(a) NH_2OH (b) NH_4NO_3
(c) N_2H_4 (d) N_3H

27. State the oxidation number of carbonyl carbon in methanal and methanoic acid respectively.

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

28. +3 oxidation state is most common in ...

(a) Ni (28) (b) Fe (26)
(c) Zn (30) (d) Cu (29)

29. In the reaction, $4Na + O_2 \longrightarrow 2Na_2O$, sodium acts as a/an

(a) oxidising agent (b) reducing agent
(c) complexing agent (d) None of these

30. The value of oxidation numbers of Cl, in Cl_2, NaOCl and ClO_3^- are respectively.

(a) +2, 0, +5 (b) 0, +2, +5
(c) +2, +3, +5 (d) 0, +1, +5

31. What is the average oxidation number of carbon in carbon suboxide?

(a) $+\frac{4}{3}$ (b) $+\frac{10}{4}$ (c) +2 (d) $+\frac{2}{3}$

32. The oxidation states of sulphur in H_2SO_4, HSO_3^- and SO_2Cl_2 respectively are

(a) +6, +4, +6 (b) +6, +6, +4
(c) +6, −6, +4 (d) −4, +6, +6

33. The largest oxidation number exhibited by an element depends on its outer electronic configuration. With which of the following outer electronic configurations the element will exhibit largest oxidation number? ***(NCERT Exemplar)***

(a) $3d^14s^2$ (b) $3d^34s^2$
(c) $3d^54s^1$ (d) $3d^54s^2$

34. Magnesium reacts with an element (X) to form an ionic compound. If the ground state electronic configuration of (X) is $1s^2\ 2s^2\ 2p^3$, the simplest formula for this compound is

(a) Mg_2X (b) MgX_2
(c) Mg_2X_3 (d) Mg_3X_2

35. The valency of Cr in $[Cr(H_2O)_4Cl_2]^+$ ion is

(a) 3 (b) 1
(c) 6 (d) 5

36. Match the Column I with Column II and select the correct option for oxidation number of N-atom from the codes given below.

	Column I (Compounds)		Column II (Oxidation number)
A.	NH_2OH	1.	−1
B.	Mg_3N_2	2.	−1
C.	N_2O	3.	+5
D.	N_2O_5	4.	−3

Codes

	A	B	C	D
(a)	1	3	4	2
(b)	2	4	3	1
(c)	2	4	1	3
(d)	4	2	1	3

Redox reactions

37. In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of
(a) +1 and +2 respectively
(b) +2 and +2 respectively
(c) +1 and +1 respectively
(d) +2 and +1 respectively

38. Which of the following arrangements represent increasing oxidation number of the central atom ? ***(NCERT Exemplar)***
(a) CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-
(b) ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-
(c) CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}
(d) CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-

39. Which of the following reactions is represented in basic medium?
(a) $MnO_4^-(aq) + SO_2(g) \longrightarrow Mn^{2+}(aq) + HSO_4^-(aq)$
(b) $H_2O_2(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + H_2O(l)$
(c) $MnO_4^-(aq) + I^-(aq) \longrightarrow MnO_2(s) + I_2(s)$
(d) $Cr_2O_7^{2-} + SO_2(g) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$

40. In the following redox reaction,
$$xUO^{2+} + Cr_2O_7^{2-} + yH^+ \longrightarrow aUO_2^{2+} + zCr^{3+} + bH_2O$$
the values of coefficients x, y and z respectively, are
(a) 3, 8, 2 (b) 3, 8, 7
(c) 3, 2, 4 (d) 3, 1, 8

41. A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction which element undergoes maximum change in oxidation number?
(a) S (b) H (c) Cl (d) C

Assertion-Reasoning MCQs

Directions In the following questions (Q.No. 42-55) a statement of Assertion followed by a statement of Reason is given. Choose the correct answer out of the following choices.
(a) Both Assertion and Reason are correct statements and Reason is the correct explanation of the Assertion.
(b) Both Assertion and Reason are correct statements, but Reason is not the correct explanation of the Assertion.
(c) Assertion is correct, but Reason is incorrect statement.
(d) Assertion is incorrect but Reason is correct statement.

42. **Assertion** $KMnO_4$ is a stronger oxidising agent than $K_2Cr_2O_7$.
Reason This is due to the increasing stability of the lower species to which they are reduced.

43. **Assertion** In the reaction between potassium permanganate and potassium iodide, permanganate ions act as oxidising agent.
Reason Oxidation state of manganese changes from +2 to +7 during the reaction.

44. **Assertion** Average oxidation number of I in KI_3 is $-1/3$.
Reason KI_3 is made up of KI and I_2. Each species have different oxidation number.

45. **Assertion** $PbCl_4$ is a powerful oxidising agent.
Reason $PbCl_4$ is more stable than $PbCl_2$.

46. **Assertion** In some cases oxygen shows positive oxidation number though it is an electronegative element.
Reason Fluorine is more electronegative than oxygen.

47. **Assertion** H_2SO_4 cannot act as a reducing agent.
Reason Sulphur cannot increase its oxidation number beyond +6.

48. **Assertion** Among halogens fluorine is the best oxidant.

REDOX reactions

Reason Fluorine is the most electronegative atom. *(NCERT Exemplar)*

49. **Assertion** The two Fe atoms in Fe_3O_4 have different oxidation numbers.

Reason Fe^{2+} ions decolourise $KMnO_4$ solution.

50. **Assertion** In the species, Br_3O_8 each of two extreme bromine exhibits oxidation state of $+6$ and the middle bromine of $+4$.

Reason The average of three oxidation numbers of bromine of the Br_3O_8 is $16/3$.

51. **Assertion** Amongst the halogens, fluorine cannot oxidise the element to the highest oxidation state.

Reason Due to small size of fluoride ion, it is difficult to oxidise flouride ion to fluorine. Hence, reverse reaction takes place more easily.

52. **Assertion** The oxidation number of O in O_3 is zero and the oxidation number of S in SO_3 is $+4$.

Reason O_3 can act as an oxidising agent as well as a reducing agent but SO_2 can act only as an oxidising agent.

53. **Assertion** The formal oxidation number of sulphur in $Na_2S_4O_6$ is 2.5.

Reason Two S-atoms are not directly linked with O-atoms.

54. **Assertion** MnO_4^- is reduced to Mn^{2+} in acidic medium.

Reason In acidic medium, following reaction takes place.

$$KMnO_4 + H_2O \longrightarrow MnO_2$$

55. **Assertion** Cl_2 gas bleaches the articles permanently.

Reason Cl_2 is a strong reducing agent.

Case Based MCQs

56. Read the passage given below and answer the following questions :

The oxidation state of an individual atom is 0. The total oxidation state of all atoms in a neutral species is 0 and in an ion is equal to the ion charge. Group 1 metals have an oxidation state of $+1$ and group 2 an oxidation state of $+2$.

The oxidation state of fluorine is -1 in compounds. Hydrogen generally has an oxidation state of $+1$ in compounds. Oxygen generally has an oxidation state of -2 in compounds.

In binary metal compounds, group 17 elements have an oxidation state of -1, group 16 elements of -2, and group 15 elements of -3. The sum of the oxidation states is equal to zero for neutral compounds and equal to the charge for polyatomic ion species. An atom is oxidised if its oxidation number increases and an atom is reduced if its oxidation number decreases.

The atom that is oxidised is the reducing agent and the atom that is reduced is the oxidising agent.

(**Note** the oxidising and reducing agents can be the same element or compound).

Redox reactions are comprised of two parts, a reduced half and an oxidised half, that always occur together.

The reduced half gains electrons and the oxidation number decreases, while the oxidised half loses electrons and the oxidation number increases.

The ion or molecule that accepts electrons is called the oxidising agent, by accepting electrons it causes the oxidation of another species. conversely, the species that donates electrons is called the **reducing agent**; when the reaction occurs, it reduces the other species.

REDOX reactions

The following questions (i-iv) are multiple choice questions. Choose the most appropriate answer :

(i) One mole of acidified $K_2Cr_2O_7$ on reaction with excess KI will liberate n mole of I_2, then the value of n is

(a) 6 (b) 1
(c) 3 (d) 7

(ii) When electrons are transferred from Zn to Cu^{2+} in copper sulphate solution, the energy (heat) is

(a) absorbed
(b) evolved
(c) consumed
(d) Both (a) and (b)

(iii) Negative $E^{\ominus}$ indicates that redox couple is

(a) weaker reducing agent than H^+/H_2 couple
(b) stronger reducing agent than H^+/H_2 couple
(c) sronger oxidising agent than H^+/H_2 couple
(d) weaker oxidising agent than H^+/H_2 couple

(iv) Which of the following statements is/are incorrect?

(a) The reactants, which undergo oxidation and reduction are called reductant and oxidant respectively
(b) In redox reaction, the oxidation number of oxidant increases, while that of reductant decreases
(c) HNO_2 acts as an oxidising as well as reducing agent
(d) Oxidation is the process, in which electrons are lost

Or

In alkaline medium, ClO_2 oxidises H_2O_2 to O_2 and itself gets reduced to Cl^-. How many moles of H_2O_2 are oxidised by 1 mole of ClO_2?

(a) 1 (b) 1.5
(c) 2.5 (d) 5

57. Read the passage given below and answer the following questions :

The concept of electron transfer is found unable to explain the redox changes or electron shift in case of covalent compounds.

To explain these changes a new concept, called oxidation number is introduced. Oxidation number is defined as the charge that an atom of the element has in its ion or appear to have when present in the combined state with other atoms.

In other words, it is also defined as the charge that an atom appear to have in a compound when all other atoms are removed as ions from the compound.

The following steps are involved while calculating the oxidation number of an atom in a given compound/ion.

***Step* I** Write down the formula of given compound/ion leaving some space between the atoms.

***Step* II** Write the oxidation state of each element above its atoms. Write down x above the atom, oxidation state of which we have to find out.

***Step* III** Multiply the oxidation numbers of each element with the number of atoms of that element present in the compound.

Enclose the product in a bracket.

***Step* IV** Equate the algebraic sum of the oxidation numbers of all the atoms present in a compound to zero or to the charge in case of ionic species charge on the ion.

***Step* V** Solve the equation obtained for the value of x.

The following questions (i-iv) are multiple choice questions. Choose the most appropriate answer:

(i) Highest oxidation state of Mn is present in

(a) $KMnO_4$ (b) K_2MnO_4
(c) Mn_2O_3 (d) MnO_2

(ii) Identify the element which never has positive oxidation number in any of its compound?

(a) Oxygen
(b) Chlorine
(c) Fluorine
(d) Bromine

REDOX reactions

(iii) When a manganous salt is fused with a mixture of KNO_3 and solid NaOH, the oxidation number of Mn changes, from + 2 to

(a) + 4 (b) + 3
(c) + 6 (d) + 7

(iv) The brown ring complex compound is formulated as $[Fe(H_2O)_5NO]SO_4$. What will be the oxidation state of iron in the given complex?

(a) + 2 (b) + 3
(c) + 4 (d) + 1

Or

In which of the following reactions, there is no change in valency?

(a) $SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$
(b) $2Na + O_2 \longrightarrow 2Na_2O_2$
(c) $Na_2O_2 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O_2$
(d) $4KClO_3 \longrightarrow 3KClO_4 + KCl$

58. Read the passage given below and answer the following questions :

In a redox reaction, the substance which oxidises the other or which itself undergoes reduction is called the oxidising agent and the substance that reduces the other and itself undergoes oxidation is called the reducing agent. e.g.

$$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$$

Thus, in the above example, $HgCl_2$ is an oxidising agent (as it reduces) and $SnCl_2$ is a reducing agent (as it oxidises).

In these questions (i-iv) a statement of Assertion followed by a statement of Reason is given. Choose the correct answer out of the following choices :

(a) Assertion and Reason both are correct statements and Reason is correct explanation for Assertion.
(b) Assertion and Reason both are correct statements but Reason is not correct explanation for Assertion.
(c) Assertion is correct statement but Reason is incorrect statement.
(d) Assertion is incorrect statement but Reason is correct statement.

(i) **Assertion** This reaction is redox reaction,

$$K_2Cr_2O_7(aq) + 3SO_2(g) + H_2SO_4(aq) \longrightarrow Cr_2(SO_4)_3(aq) + K_2SO_4(aq) + H_2O(l)$$

Reason Both oxidation and reduction take place in the given reaction.

(ii) **Assertion** The decomposition of hydrogen peroxide to form water and oxygen is an example of disproportionation reaction.

Reason The oxygen of peroxide is in -1 oxidation state and it is converted to zero oxidation state in O_2 and -2 oxidation state in H_2O.

(iii) **Assertion** In the species, Br_3O_8 each of two extreme bromine exhibits oxidation state of $+6$ and the middle bromine of $+4$.

Reason The average of three oxidation numbers of bromine of the Br_3O_8 is 16/3.

(iv) **Assertion** In the reaction between potassium permanganate and potassium iodide, permanganate ions act as oxidising agent.

Reason Oxidation state of manganese changes from +2 to +7 during the reaction.

Or

Assertion A negative value of $E°$ means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.

Reason A negative $E°$ means that the redox couple is stronger reducing agent than the H^+/H_2.

59. Read the passage given below and answer the following questions :

The real or imaginary charge which an atom appears to have in its combined state is called oxidation state or oxidation number of that atom. Fraction oxidation states are often used to represent the average oxidation states of several atom in a structure. These oxidation states are very helpful in finding the oxidation and reduction process in redox reactions. Redox reactions are of two main, i.e intermolecular redox reactions and intramolecular redox reactions.

The elements that show an increase in oxidation number (hydrogen and chlorine in the above reaction) are **oxidised**, while the elements that are reduced (oxygen and chlorine in the above reaction) show a decrease in their oxidation numbers from their initial values.

In these questions (i-iv) a statement of Assertion followed by a statement of Reason is given. Choose the correct answer out of the following choices :

(i) **Assertion** Oxidation state of nitrogen in N_3H is $-\frac{1}{3}$.

Reason Nitrogen is less electronegative than hydrogen.

(ii) **Assertion** The decomposition of hydrogen peroxide to form water and oxygen is an example of disproportionation reaction.

Reason The oxygen of peroxide is in -1 oxidation state and it is converted to zero oxidation state in O_2 and -2 oxidation state in H_2O.

(iii) **Assertion** The electrons are transferred from zinc to copper through the wire which connects the two rods.

Reason Electricity flows through the salt-bridge by migration of ions from one beaker to other.

(iv) **Assertion** Redox couple is the combination of oxidised and reduced, form of a substance involved in an oxidation or reduction half-cell.

Reason In the representation $E^{\ominus}_{Fe^{3+}/Fe^{2+}}$ and $E^{\ominus}_{Cu^{2+}/Cu}$, Fe^{3+}/Fe^{2+} and Cu^{2+}/Cu are redox couples.

Or

Assertion Oxidation number of hydrogen -1 in CaH_2.

Reasons CaH_2 is a metal hydrides and for hydrides, hydrogen is assigned the oxidation number of -1.

ANSWERS

Multiple Choice Questions

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

Assertion-Reasoning MCQs

36. (a) 37. (a) 38. (c) 39. (b) 40. (a) 41. (d) 42. (d) 43. (c) 44. (c) 45. (c)
46. (d) 47. (b) 48. (d) 49. (c) 50. (d)

Case Based MCQs

51. (i) - (c); (ii) - (d); (iii) - (a); (iv) - (b); (v) - (c)
52. (i) - (a); (ii) - (c); (iii) - (d); (iv) - (b); (v) - (b)
53. (i) - (c); (ii) - (a); (iii) - (a); (iv) - (d); (v) - (c)
54. (i) - (b); (ii) - (a); (iii) - (c); (iv) - (a); (v) - (c)

REDOX reactions EXPLANATIONS

1. Electrochemical processes for the extraction of highly reactive metals and non-metals, manufacturing of chemical compounds like caustic soda, operation of dry and wet batteries and corrosion of metals fall within the range of redox processes.

2. Addition of oxygen/electronegative element and removal of hydrogen/electropositive element takes place in oxidation.

3. Reduction is a process which involves addition of hydrogen or electropositive elements to a substance or removal of oxygen or electronegative element from a substance.

4. In the given reaction, removal of electronegative element, i.e. chlorine from ferric chloride takes place. Hence, it is an example of reduction process.

$$2FeCl_3(aq) + H_2(g) \longrightarrow 2FeCl_2(aq) + 2HCl(aq)$$

Removal of chlorine (electronegative element)

5. $H_2S + Cl_2 \longrightarrow 2HCl + S$

Addition of hydrogen (reduction)

Removal of hydrogen (oxidation)

Thus, H_2S is oxidised and Cl_2 is reduced.

6. In the given reaction,

$$2K_4[Fe(CN)_6](aq) + H_2O_2(aq) \longrightarrow 2K_3[Fe(CN)_6](aq) + 2KOH(aq)$$

Removal of potassium (electropositive element)

So, here oxidation takes place due to removal of one potassium atom.

7. Oxidation is a process, which involves addition of oxygen/electronegative element to a substance or removal of hydrogen/electropositive element from a substance.

e.g. $2Mg + O_2 \longrightarrow 2MgO$

Addition of oxygen = Oxidation

8. $CH_2{=}CH_2 + H_2 \longrightarrow H_3C-CH_3$

(Addition of hydrogen)

Reduction of ethylene occurs due to the addition of hydrogen.

9. In oxidation process, oxidation number increases and number of electrons decreases.

10. $2\overset{+3}{Co}(OH)_3 + \overset{0}{Cd} + 2H_2O \longrightarrow 2\overset{+2}{Co}(OH)_2 + \overset{+2}{Cd}(OH)_2$

Increase in ON

Here, oxidation number of Cd increases from 0 to +2, hence Cd is oxidised.

11. Both oxidation and reduction are taking place in the following reaction,

Oxidised (reducing agent)

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

Reduced (oxidising agent)

H_2-reducing agent; Br_2-oxidising agent.

In other options, neither oxidation nor reduction takes place because oxidation number of elements involved in reaction remain same.

12. $2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \longrightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}$

Loss of e^- (oxidation)

Gain of e^- (reduction)

13. Oxygen in H_2O_2 has oxidation number -1 which can increase or decrease.

Hence, it can act as both oxidising agent or reducing agent.

14. Hydrogen is oxidised by loss of 1 electron

$$2\overset{0}{H_2} + \overset{0}{O_2} \longrightarrow 2\overset{+1}{H_2}\overset{-2}{O}$$

Oxygen is reduced by gain of 2 electrons

In this reaction, hydrogen (H) has transferred electrons to oxygen (O).

15. Following are the examples of redox reaction :

(a) $CuO + H_2 \longrightarrow Cu + H_2O$

(b) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

(c) $2K + F_2 \longrightarrow 2KF$

Option (d) is not an example of redox reaction.

redox reactions

16. The correct match is

A → 3, B → 4, C → 1, D → 2

Oxidation (addition of oxygen)

(A) $2Mg + O_2 \longrightarrow 2MgO$

Oxidation (addition of electronegative element)

(B) $Mg + Cl_2 \longrightarrow MgCl_2$

(C) $2H_2S + O_2 \longrightarrow 2S + 2H_2O$

Reduction (removal of hydrogen)

(D) $2KI + H_2O + O_3 \longrightarrow 2KOH + I_2 + O_2$

Oxidation (removal of electropositive element)

17. Since, K contains only one electron in its outermost shell, it has higher tendency to donate it, i.e. has higher tendency to get oxidised.

Therefore, it is the strongest reducing agent among the given elements.

18. $4Na + O_2 \longrightarrow 2Na_2O$

Loss of e^- (oxidation)

In the above reaction, Na converts into (Na^+) ion, i.e. Na donates its electron to oxygen atom. So, it behaves as reducing agent.

19. $C_6H_6 + H\overset{+5}{N}O_3 \longrightarrow C_6H_5 - \overset{+3}{N}O_2 + H_2O$

Reduction (gain of e^-)

In this reaction, HNO_3 behaves as an oxidising agent while in rest of the reactions such as,

(i) $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

(ii) $Ca(OH)_2 + 2HNO_3 \longrightarrow Ca(NO_3)_2 + 2H_2O$

(iii) $NaCl + HNO_3 \longrightarrow HCl + NaNO_3$

HNO_3 neither behaves as oxidising agent nor as reducing agent.

20.

Increase in ON (oxidation +2)

$$14H^+ + \overset{+6}{Cr_2}O_7^{2-} + 3Ni \longrightarrow 2Cr^{3+} + 7H_2O + 3\overset{+2}{Ni}$$

Decrease in ON (reduction)

As in this reaction Ni is oxidised to Ni^{2+} and reduced to $Cr_2O_7^{2-}$, thus it acts as reducing agent.

Loss of e^- (oxidation)

21. $H_2O + \overset{0}{Br_2} \longrightarrow \overset{+}{H}\,\overset{-}{Br} + H\overset{-}{O}\overset{+}{Br}$

Gain of e^- (reduction)

Here, oxidation number of bromine increases as well as decreases, i.e. bromine is oxidised as well as reduced.

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

(a) $\overset{x}{N_2}\overset{+1}{H_4}$, $2x + (+1)\,4 = 0$

$$2x = -4$$
$$x = -2$$

(b) $\overset{x}{N}\overset{+1}{H_3}$, $x + (+1)\,3 = 0$

$$x = -3$$

(c) $\overset{x}{N_3}\overset{+1}{H}$, $3x + (+1) = 0$

$$3x = -1$$
$$x = -1/3$$

(d) $\overset{x}{N}\overset{+1}{H_2}\overset{-2}{O}\overset{+1}{H}$,

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

The oxidation state of nitrogen is highest in N_3H.

23. We know that, the algebraic sum of the oxidation states is always zero in neutral compound.

Oxidation states of $X = +2$

$Y = +5$

$Z = -2$

So, the algebraic sum of total X, Y and Z should be equal to zero which is found in XY_2Z_6.

$$XY_2Z_6 = +2 + (5 \times 2) + (-2 \times 6)$$
$$= +2 + 10 - 12 = 0$$

24. (a) $K_4Fe(CN)_6 \Rightarrow 1 \times 4 + x + (-1)6 \Rightarrow 0$; $\therefore x = 2$

(b) $Fe_3O_4 \Rightarrow FeO.Fe_2O_3 \Rightarrow +2, +3$

(c) $Fe(CO)_5 \Rightarrow x + 0(5) = 0$; $\therefore x = 0$

(d) $FeSO_4(NH_4)_2SO_4.6H_2O \Rightarrow x + (-2) + 0 + 0 = 0$

$\therefore \quad x = +2$

Thus, in $Fe(CO)_5$, Fe shows minimum oxidation number.

25. (a) $Na_2\overset{2.5}{S_4}O_6$; $H_2\overset{+6}{S_2}O_7$; $Na_2\overset{+2}{S_2}O_3$; $\overset{0}{S_8}$

(b) $H_2\overset{+6}{S}O_4$; $\overset{+4}{S}O_2$; $H_2\overset{-2}{S}$; $H_2\overset{+6}{S_2}O_8$

(c) $\overset{+2}{S}O_2^2$; $\overset{+6}{S}O_4^2$; $\overset{+4}{S}O_3^{2-}$; $H\overset{+6}{S}O_4^-$

(d) $H_2\overset{+6}{S}O_5$; $H_2\overset{+4}{S}O_3$; $\overset{+2}{S}Cl_2$; $H_2\overset{-2}{S}$

Thus, in option (d), compounds are arranged in the decreasing order of oxidation number of sulphur.

REDOX reactions

26. NH_4NO_3 is actually NH_4^+ and NO_3^-. It is an ionic compound.

The oxidation number of nitrogen in the two species is different as shown below:

In NH_4^+,

$\Rightarrow x+(4+1)=+1$ or $x+4=+1$

or $x=-3$

Let, oxidation number of N in NO_3^- is x

$\Rightarrow x+(3x-2)=-1$ or $x-6=-1$ or $x=+1$

27. Let the oxidation number of carbonyl carbon in methanal (HCHO) and methanoic acid (HCOOH) is x and y respectively.

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

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

In HCOOH, $2(+1)+y+2(-2)=0$

$2+y-4=0 \Rightarrow y=+2$

28. (a) $Ni(III) = 3d^7 4s^0 = \underbrace{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}_{3d}\ \underbrace{\boxed{}}_{4s}$

(b) $Fe(III) = 3d^5 4s^0 = \underbrace{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}_{3d}\ \underbrace{\boxed{}}_{4s}$

$+3$ oxidation state of Fe provides the extra stability due to half-filled d-orbitals.

(c) $Zn(III) = 3d^9 4s^0 = \underbrace{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}}_{3d}\ \underbrace{\boxed{}}_{4s}$

(d) $Cu(III) = 3d^8 4s^0 = \underbrace{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}}_{3d}\ \underbrace{\boxed{}}_{4s}$

29. $4\overset{(0)}{Na}+O_2 \longrightarrow 2\overset{+1}{Na}_2O$

Loss of e^- (oxidation)

In this reaction,

Na converts into (Na^+) ion, i.e. Na donates its electron to oxygen atom. So, it behaves as an reducing agent.

30. Oxidation number of Cl in molecular state (i.e. in Cl_2) is zero.

Let, oxidation number of Cl in NaOCl is x.

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

$x=+1$

Let, oxidation number of Cl in ClO_3^-.

$\therefore \quad x+3(-2)=-1$

$x=+5$

31. $O=\overset{+2}{C}=\overset{0}{C}=\overset{+2}{C}=O$

Carbon suboxide

In C_3O_2, two C-atoms linked with oxygen atoms are present in +2 oxidation state and central carbon has zero oxidation state.

So, the average oxidation state of carbon is $+\frac{4}{3}$.

32. (i) Let oxidation number of S in H_2SO_4 is x.

$\therefore +1\times2+x+(-2)\times4=0$

$x=8-2$ or $x=+6$

Therefore, oxidation number of S in H_2SO_4 is +6.

(ii) Let oxidation number of S in HSO_3^- is x.

$(+1)+x+(-2)3=-1$

$x=-1+5$ or $x=+4$

Therefore, oxidation number of S in HSO_3^- is +4.

(iii) Let the oxidation number of S in SO_2Cl_2 be x.

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

$x=+6$

Therefore, in SO_2Cl_2 the oxidation number of sulphur is +6

33. Highest oxidation number of any transition element = $(n-1)$ d electrons + ns electrons. Therefore, larger the number of electrons in the $3d$-orbitals, higher is the maximum oxidation number.

(a) $3d^1 4s^2 = 3$

(b) $3d^3 4s^2 = 3+2=5$

(c) $3d^5 4s^1 = 5+1=6$ and

(d) $3d^5 4s^2 = 5+2=7$

Thus, option (d) is correct.

34. Given, electronic configuration of $X=1s^2 2s^2 2p^3$

$\therefore$ The valency of X will be 3.

The valency of Mg is +2.

$\therefore$ Magnesium reacts with element X to form an ionic compound with formula Mg_3X_2.

35. The valency of Cr in $[Cr(H_2O)_4Cl_2]^+$ ion is x.

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

$x=1+2$

$x=3$

So, the valency of Cr in $[Cr(H_2O)_4Cl_2]^+$ ion is 3.

REDOX reactions

36. The correct match is
$A \rightarrow 2, B \rightarrow 4, C \rightarrow 1, D \rightarrow 3.$
The oxidation number of N-atom in given compounds are shown below :

$$\begin{array}{cccc} \overset{-1+1-2+1}{NH_2OH}, & \overset{+2\ \ -3}{Mg_3N_2} & \overset{+1\ -2}{N_2O} & \overset{+5\ -2}{N_2O_5} \\ \downarrow & \downarrow & \downarrow & \downarrow \\ -1 & -3 & -1 & +5 \end{array}$$

37. Electronegativity of fluorine is more than that of oxygen atom, so F gains electron with negative charge.
In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), oxygen transfers electron to fluorine atom. Thus,
Oxidation number of oxygen in $OF_2 = +2.$
Oxidation number of oxygen in $O_2F_2 = +1.$

38. Writing the oxidation number (O.N.) of Cr, Cl and Mn on each species in the four set of ions, then,

(a) $\overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^-, \overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-$

(b) $\overset{+5}{Cl}O_3^-, \overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-, \overset{+3}{Cr}O_2^-$

(c) $\overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^-, \overset{+7}{Mn}O_4^-, \overset{+6}{Cr}O_4^{2-}$

(d) $\overset{+6}{Cr}O_4^{2-}, \overset{+7}{Mn}O_4^-, \overset{+3}{Cr}O_2^-, \overset{+5}{Cl}O_3^{3-}$

Only in the arrangement (a), the ON of central atom increases from left to right, therefore, option (a) is correct.

39. $\overset{+7}{Mn}O_4^-(aq) + \overset{-1}{I^-}(aq) \longrightarrow \overset{+4}{MnO_2}(s) + \overset{0}{I_2}(s)$

This reaction is represented in basic medium because in basic medium $\overset{+7}{Mn}O_4^-$ is reduced to $\overset{+4}{Mn}O_2$ (i.e. $\overset{+7}{Mn}$ to $\overset{+4}{Mn}$), while in acidic medium, MnO_4^- is reduced from Mn^{7+} to Mn^{2+}.

40. The balanced chemical reaction is given as

$$3UO^{2+} + Cr_2O_7^{2-} + 8H^+ \longrightarrow 3UO_2^{2+} + 2Cr^{3+} + 4H_2O$$

Hence, the value of x, y and z are respectively 3, 8 and 2.

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

$$\overset{+1+5-2}{KClO_3} + \overset{+1+3-2}{H_2C_2O_4} + \overset{+1+6-2}{H_2SO_4} \xrightarrow{\Delta} \overset{+1+6-2}{K_2SO_4} + \overset{+1-1}{KCl} + \overset{+4-2}{CO_2} + \overset{+1-2}{H_2O}$$

Thus, Cl is the element which undergoes maximum change in the oxidation state from +5 to −1.

42. Both Assertion and Reason are correct explanation and Reason is the correct explanation for Assertion.

43. The reaction of potassium permanganate and potassium iodide is as follows :

$$10KI + 2K\overset{+7}{Mn}O_4 + 8H_2SO_4 \longrightarrow 2\overset{+2}{Mn}SO_4 + 6K_2SO_4 + 8H_2O + 5I_2$$

Oxidation state of Mn decreases from +7 to +2.
Thus, Assertion is correct but Reason is incorrect.

44. KI_3 dissociates into KI and I_2.
$\therefore$ Average oxidation number of I

$$= \frac{-1\times1 + 0\times2}{1+2} = \frac{-1}{3}$$

Thus, either two values are reported separately or one value is reported.

45. $PbCl_2$ is more stable than $PbCl_4$ or Pb^{2+} is more stable than Pb^{4+}(due to the inert pair effect)

$$2e^- + \underset{\text{(Oxidising agent)}}{Pb^{4+}} \longrightarrow Pb^{2+} \text{ (Reduction)}$$

Assertion is correct but Reason is incorrect statement.

46. Oxygen is the most electronegative element after fluorine. Therefore, in the compounds between oxygen and fluorine, oxygen is found to show positive oxidation state.
Both Assertion and Reason are correct statements and Reason is the correct explanation of the Assertion.

47. Maximum oxidation state of S is + 6, it cannot exceed beyond, it. Therefore it, cannot be further oxidised.
Both Assertion and Reason are correct statements and Reason is the correct explanation of the Assertion.

48. Among halogen F_2 is the best oxidant because it has the highest $E°$ value.
Both Assertion and Reason are correct, but Reason is not the correct explanation of Assertion.

49. $Fe_3O_4 \equiv (\overset{+2}{Fe}O \cdot \overset{+2}{Fe_2}O_3)$

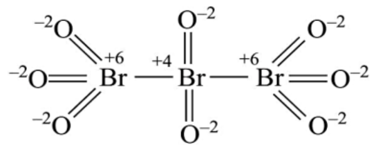
The oxidation states of Fe in FeO and Fe_2O_3 are + 2 and + 3.

$$Fe^{2+} + \underset{\text{(Pink)}}{MnO_4^-} \longrightarrow Fe^{3+} + \underset{\text{(Colourless)}}{Mn^{2+}}$$

REDOX reactions

Both Assertion and Reason are correct statements, but Reason is not the correct explanation of the Assertion.

50. The structure of Br_3O_8 (tribromooctaoxide) is



Thus, oxidation state of two corner Br atoms is +6 and of middle one is +4. The difference in oxidation states is due to difference in bonding situations.

Average oxidation state $= \dfrac{+6+4+6}{3} = \dfrac{16}{3}$

Thus, both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

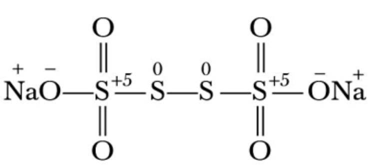
51. Due to high electronegativity and high heat of association, fluorine oxidises the elements to their highest oxidation state.

Both Assertion and Reason are correct statements, but Reason is not the correct explanation of the Assertion.

52. SO_2 can act both as an oxidising and a reducing agent. O_3 can act only as an oxidising agent.

The oxidation number of O in O_3 is zero. It can only decrease from zero to -1 or -2 but cannot increase to $+2$. Therefore, it can act as an oxidising agent only. In SO_2, the oxidation number of S is $+4$. It can have a minimum oxidation number of -2 and maximum of $+6$. Its oxidation number either decreases or increases and hence, it can act both as an oxidising and a reducing agent.

Assertion is correct, but Reason is incorrect statement.

53. 

Formal oxidation number of sulphur

$$= \frac{2\times 5 + 2\times 0}{4} = 2.5$$

Assertion is correct, but Reason is incorrect statement.

54. MnO_4^- is reduced to MnO_2 in a mild basic medium or neutral medium whereas in an acidic medium,

MnO_4^-, is reduced to Mn^{2+} and in a strong basic medium, it is reduced to MnO_4^{2-}.

So, Assertion is incorrect but Reason is correct.

55. Cl_2 is an oxidising agent. It bleaches the articles permanently by oxidation in presence of moisture.

Assertion is correct, but Reason is incorrect statement.

56. (i) $K_2Cr_2O_7 + \text{Excess KI} \longrightarrow Cr^{3+} + H_2O + nI_2^0$

$$\overset{6+}{Cr} \xrightarrow{+3e^-} \overset{3+}{Cr}$$

Cr^{6+} accepts $3e^-$, so mole of $I_2 = 3$.

(ii) If zinc rod is dipped in copper sulphate solution, then due to transfer of electron from zinc to copper ion, heat is evolved.

(iii) A negative $E^\ominus$ means that the redox couple is a stronger reducing agent than H^+/H_2 couple.

$-E^\ominus$ = strong reducing agent

$+E^\ominus$ = weak reducing agent

(iv) Oxidant = oxidising agent = reduces itself by increasing oxidation number.

Reductant = reducing agent = oxidising itself, by decreasing oxidation number.

Or

In alkaline medium, the balanced equation is,

$$2ClO_2 + 5H_2O_2 + 2OH^- \longrightarrow 2Cl^- + 5O_2 + 6H_2O$$

2 moles of ClO_2 oxidise 5 moles of H_2O_2

$\therefore$ 1 mole of ClO_2 will oxidise $= \dfrac{5}{2}\times 1 = 2.5.$

57. (i) Let the oxidation state of Mn be x.

(a) **$KMnO_4$,**

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

(b) **K_2MnO_4,**

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

(c) **Mn_2O_3,**

$$x \times 2 + (-2)\times 3 = 0$$
$$2x = +6$$
$$x = +3$$

REDOX reactions

(d) $\mathbf{MnO_2}$, $x + (-2) \times 2 = 0$

$$x = 4$$

Thus, oxidation state of Mn is highest in $KMnO_4$.

(ii) Fluorine is the most electronegative element.

Hence, does not possess positive oxidation number in any of its compound.

(iii) $\overset{(+2)}{Mn^{2+}} + 2NO_3^- + 2OH^- \longrightarrow$

$$\overset{(+6)}{MnO_4^{2-}} + 2H_2O + 2NO$$

The oxidation number of Mn changes from +2 to + 6.

(iv) Let oxidation state of Fe in complex is x.

$\therefore \quad x + 5(0) + (-1) = -2$

or $\quad x = +3$

Or

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

(b) $2\overset{0}{Na} + \overset{0}{O_2} \longrightarrow 2\overset{+1}{Na_2}\overset{-1}{O_2}$

(c) $2\overset{+1}{Na_2}\overset{-1}{O_2} + H_2SO_4 \longrightarrow \overset{+1}{Na_2}SO_4 + H_2\overset{-1}{O_2}$

(no change in valency)

(d) $4K\overset{+5}{Cl}O_3 \longrightarrow 3K\overset{+7}{Cl}O_4 + K\overset{-1}{Cl}$

58. (i) The redox change in the given reaction is as follows :

Reduction

$$\underset{(+6)}{K_2Cr_2O_7} + \underset{(+4)}{3SO_2} + H_2SO_4 \longrightarrow \underset{(+3)\ (+6)}{Cr_2(SO_4)_3}$$

Oxidation

$$+ \underset{(+6)}{K_2SO_4} + H_2O$$

$\therefore$ Both Assertion and Reason are correct and Reason is the correct explanation of Assertion.

(ii)

Oxidation

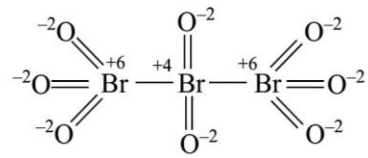
$$2H_2\overset{-1}{O_2} \longrightarrow 2H_2\overset{-2}{O} + \overset{0}{O_2}$$

Reduction

Thus, the above reaction is an example of disproportionation reaction.

Thus, both Assertion and Reason are correct and Reason is the correct explanation of Assertion.

(iii) The structure of Br_3O_8 (tribromooctaoxide) is



Thus, oxidation state of two corner Br atoms is +6 and of middle one is +4. The difference in oxidation states is due to difference in bonding situations.

Average oxidation state $= \dfrac{+6+4+6}{3} = \dfrac{16}{3}$

Thus, both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

(iv) The reaction of potassium permanganate and potassium iodide is as follows :

$$10KI + 2K\overset{+7}{Mn}O_4 + 8H_2SO_4 \longrightarrow$$

$$2\overset{+2}{Mn}SO_4 + 6K_2SO_4 + 8H_2O + 5I_2$$

Oxidation state of Mn decreases from +7 to +2.

Thus, Assertion is correct but Reason is incorrect.

Or

As we know H^+/H_2 couple has zero standard reduction potential so, ions having positive $E°$ value are weaker reducing agent, while ions having negative $E°$ value are stronger reducing agent. Thus, Assertion is incorrect but Reason is correct.

59. (i) Let, oxidation state of N in N_3H be x.

$$3x + 1 = 0$$

$$x = -\frac{1}{3}$$

Nitrogen is more electronegative than hydrogen.

$\therefore$ Assertion is correct but Reason is incorrect.

(ii)

Oxidation

$$2H_2\overset{-1}{O_2} \longrightarrow 2H_2\overset{-2}{O} + \overset{0}{O_2}$$

Reduction

REDOX reactions

Thus, the above reaction is an example of disproportionation reaction.

Both Assertion and Reason are correct and Reason is the correct explanation of Assertion.

(iii) The electrons are transferred from Zn to Cu^{2+} through the metallic wire which connects the two rods.

While electricity flows through the salt-bridge by migration of ions from one beaker to other.

Both Assertion and Reason are correct statements but Reason is not the correct explanation of Assertion.

(iv) Redox couple is the combination of oxidised and reduced form of a substance.

In the representation $E^\circ_{Fe^{3+}/Fe^{2+}}$ and $E^\circ_{Cu^{2+}/Cu}$, Fe^{3+}/Fe^{2+} and Cu^{2+}/Cu are redox couples.

Both Assertion and Reason are correct statements but Reason is not the correct explanation of Assertion.

Or

Oxidation number of elements in their compounds or ions is obtained using some rules, e.g. hydrogen is assigned oxidation number of + 1 in general and – 1 for metal hydrides.

Both Assertion and Reason are correct statements but Reason is not the correct explanation of Assertion.