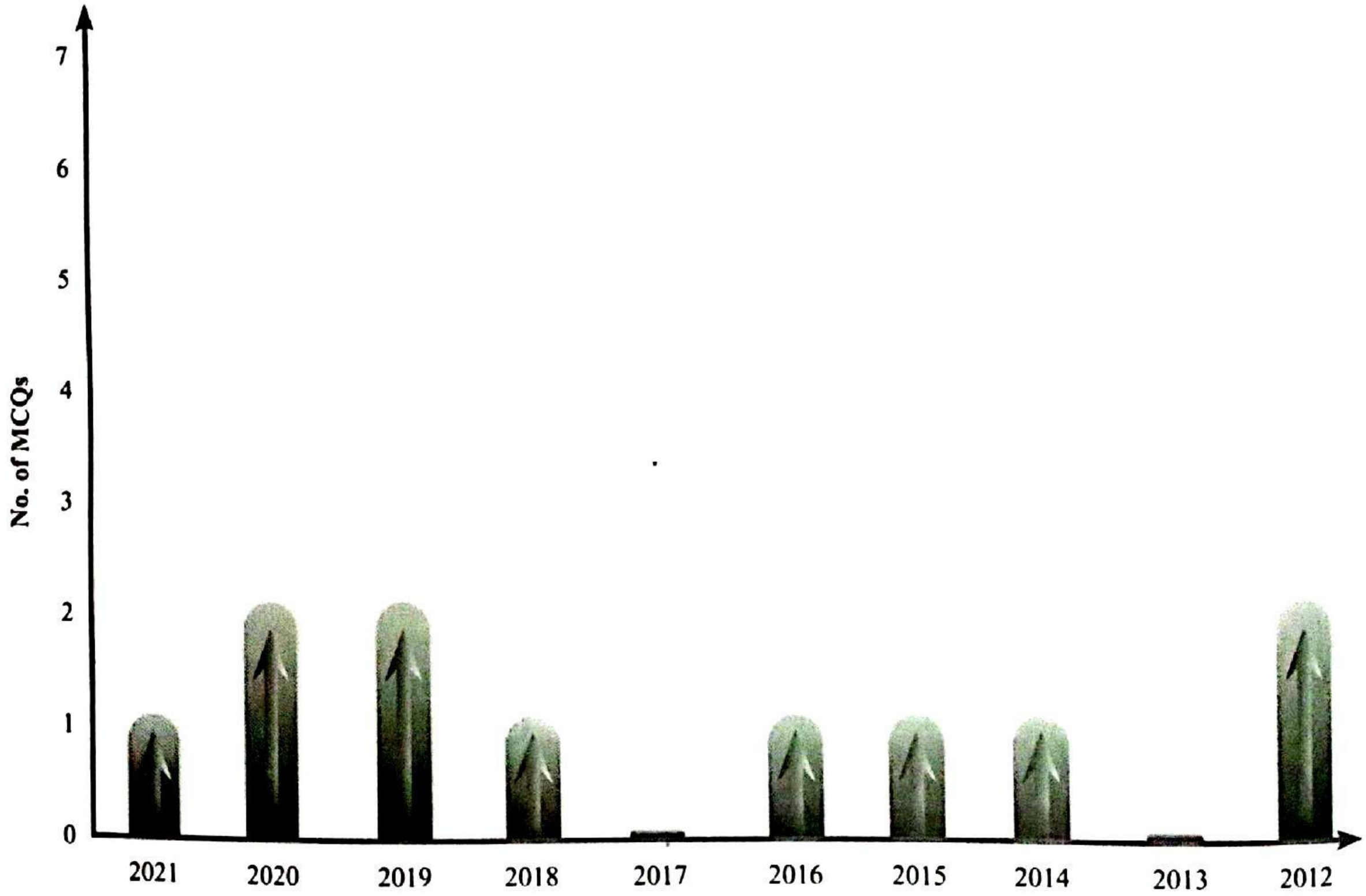
8

REDOX REACTIONS

Past Years NEET Trend



Investigation Report

TARGET EXAM	PREDICTED NO. OF MCQs	CRITICAL CONCEPTS
NEET	0-1	• Redox reaction and electrode processes • Fractional oxidation number, Balancing of redox reactions

Perfect Practice Plan

Topicwise Questions	Learning Plus	Multiconcept MCQs	NEET Past 10 Years Questions	Total MCQs
67	29	20	11	127

INTRODUCTION

Many Chemical reactions involve transfer of electrons from one chemical substance to another. These electron-transfer reactions are termed as **Oxidation-Reduction** or **Redox reactions.**

or

Those reactions which involve oxidation and reduction both simultaneously are known as oxidation reduction or Redox Reaction.

CLASSICAL IDEA OF REDOX REACTION

Oxidation

Addition of Oxygen or any other electronegative element.

E.g., $2Mg + O_2 \rightarrow 2MgO$

Removal of Hydrogen e.g., $H_2S + Cl_2 \rightarrow 2HCl + S$

Increase in positive charge e.g., $Fe^{2+} \rightarrow Fe^{3+} + e^-$

Increase in oxidation number

E.g., $\overset{(+2)}{SnCl_2} \rightarrow \overset{(+4)}{SnCl_4}$

Removal of electron e.g. $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

Reduction

Removal of Oxygen or any other electronegative element.

E.g., $CuO + C \rightarrow Cu + CO$

Addition of Hydrogen e.g. $S + H_2 \rightarrow H_2S$

Decrease in positive charge e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$

Decrease in oxidation number

E.g., $\overset{(+7)}{MnO_4^-} \rightarrow \overset{(+2)}{Mn^{2+}}$

Addition of electron e.g. $Fe^{3+} + e^- \rightarrow Fe^{2+}$

REDOX REACTION IN TERM OF ELECTRON TRANSFER REACTIONS (MODERN IDEA)

Oxidation

De-electronation.

Oxidation process is those process in which one or more e^- are lost by an atom, ion or molecule.

Example: $Zn \rightarrow Zn^{+2} + 2e^-$

$M \rightarrow M^{n+} + ne^-$

Reduction

Electronation.

Reduction process is those process in which one or more e^- are gained by an atom, ion or molecule. Example:

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

$M^{n+} + ne^- \rightarrow M$

Oxidising agent or Oxidant

Oxidising agents are those compounds which can oxidise others and reduce itself during the chemical reaction.

Those reagents in which oxidation number of an element decreases or which undergoes gain of electrons in a redox reaction are termed as oxidants.

E.g., $KMnO_4$, $K_2Cr_2O_7$, HNO_3, conc.H_2SO_4 etc are powerful oxidising agents.

Reducing agent or Reductant

Reducing agents are those compounds which can reduce other and oxidise itself during the chemical reaction.

Those reagents in which, for an element, oxidation number increases or which undergoes loss of electrons in a redox reaction are termed as reductants.

E.g., KI, $Na_2S_2O_3$ etc are the powerful reducing agents.

> ✍ **KEY NOTE**
>
> There are some compounds also which can work both as oxidising agent and reducing agent
>
> E.g., H_2O_2, NO_2^-

OXIDATION NUMBER

It is defined as an imaginary or apparent charge developed over atom of an element when it goes from its elemental free state to combined state in molecules.

It is calculated on basis of an arbitrary set of rules.

It is a relative charge in a particular bonded state.

In this method, it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.

RULES GOVERNING OXIDATION NUMBER

The following rules are helpful in calculating oxidation number of the elements in their different compounds. It is to be known that the basis of these rule is the electronegativity of the element .

Fluorine atom: Fluorine is most electronegative atom (known). It always has oxidation number equal to –1 in all its compounds

Oxygen atom: In general and as well as in its oxides , oxygen atom has oxidation number equal to –2.

In case of

(*i*) peroxide (e.g. H_2O_2, Na_2O_2) is –1,

(*ii*) super oxide (e.g. KO_2) is –1/2

(*iii*) ozonide (e.g. KO_3) is –1/3

(*iv*) OF_2 is + 2 & O_2F_2 is +1

Hydrogen atom: In general, H atom has oxidation number equal to +1. But in metallic hydrides (e.g. NaH, KH), it is –1.

Halogen atom: In general, all halogen atoms (Cl, Br, I) have oxidation number equal to –1.

But if halogen atom is attached with a more electronegative atom than halogen atom, then it will show positive oxidation numbers.

$\overset{+5}{KClO_3}, \overset{+5}{HIO_3}, \overset{+7}{HClO_4}, \overset{+5}{KBrO_3}$

Metals:

(*i*) Alkali metals (Li, Na, K, Rb, ...) always have oxidation number +1

(*ii*) Alkaline earth metals (Be, Mg, Ca .....) always have oxidation number +2.

(*iii*) Aluminium always has +3 oxidation number

> ✍ **KEY NOTE**
>
> Metal may have negative or zero oxidation number

Oxidation number of an element in free state or in allotropic forms is always zero

E.g., $\overset{0}{O_2}$, $\overset{0}{S_8}$, $\overset{0}{P_4}$, $\overset{0}{O_3}$

Sum of the oxidation numbers of atoms of all elements in a molecule is zero.

Sum of the oxidation numbers of atoms of all elements in an ion is equal to the charge on the ion .

If the group number of an element in modern periodic table is ***n***, then its oxidation number may vary from

(n – 10) to ***(n – 18)*** (but it is mainly applicable for p-block elements)

e.g. N- atom belongs to 15th group in the periodic table, therefore as per rule, its oxidation number may vary from

$$-3 \text{ to } +5 \left(\overset{-3}{N}H_3, \overset{+2}{N}O, \overset{+3}{N_2}O_3, \overset{+4}{N}O_2, \overset{+5}{N_2}O_5 \right)$$

The maximum possible oxidation number of any element in a compound is never more than the number of electrons in valence shell.(but it is mainly applicable for p-block elements)

O.S. of an atom may be fractional, negative, zero as well as positive.

TRAIN YOUR BRAIN

Q. Calculate oxidation number of underlined element:

(a) $Na_2\underline{S}_2O_3$ (b) $Na_2\underline{S}_4O_6$

Ans. (a) Let oxidation number of S-atom is x. Now work accordingly with the rules given before.

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

$x = +2$

(b) Let oxidation number of S-atom is x

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

$x = +2.5$

KEY NOTE

It is important to note here that $Na_2S_2O_3$ have two S-atoms and there are four S-atom in $Na_2S_4O_6$. However none of the sulphur atoms in both the compounds have + 2 or + 2.5 oxidation number, it is the average of oxidation number, which reside on each sulphur atom. Therefore, we should work to calculate the individual oxidation number of each sulphur atom in these compounds.

CALCULATION OF INDIVIDUAL OXIDATION NUMBER

To calculate individual oxidation number of the element in its compound we should know the structure of the compound and use the following guidelines.

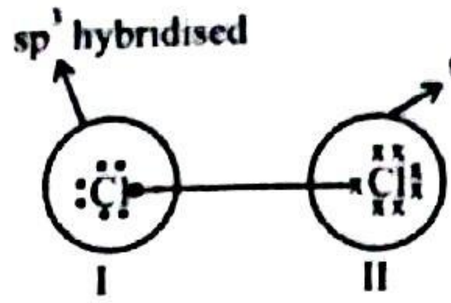
Formula: Oxidation Number = Number of e^- in the valence shell – Number of e^- taken up after bonding

Guidelines: It is based on electronegativity of elements.

1. If there is a bond between similar type of atom and each atom has same type of hybridisation, then bonded pair electrons are equally shared by each element.

Example: Oxidation number of each Cl-atom in Cl_2 molecule

Structure:



(I) Number of electrons in the valence shell = 7.

Number of electrons taken up after bonding = 7.

$\therefore$ Oxidation number = 7 – 7 = 0.

(II) similarly, oxidation number = 7 – 7 = 0

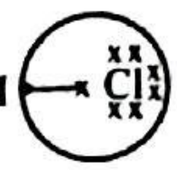
2. If there is a bond between different type of atoms:

e.g. A – B (if B is more electronegative than A)

Then after bonding, bonded pair of electrons are counted with B - atom .

Example: Oxidation number of each atom in HCl molecule

Structure:



H : Number of electrons in the valence shell = 1

Number of electrons taken up after bonding = 0

Oxidation number of H = 1 – 0 = + 1

Cl : Number of electrons in the valence shell = 7

Number of electrons taken up after bonding = 8

Oxidation number of Cl = 7– 8 = – 1

KEY NOTE

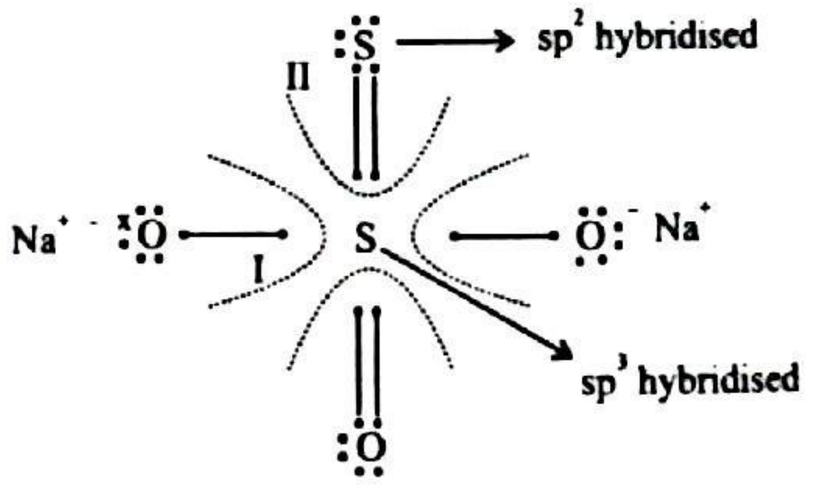
Electron of H-atom is now counted with Cl-atom, because Cl-atom is more electronegative than H-atom

TRAIN YOUR BRAIN

Q. Calculate individual oxidation number of each S-atom in $Na_2S_2O_3$ (sodium thiosulphate) with the help of its structure .

Ans.

Structure:



I (central S-atom) is sp^3 hybridised (25% s-character) and II (terminal S-atom) is sp^2 hybridised (33% s-character). Therefore, terminal sulphur atom is more electronegative than central sulphur atom. Now, the shared pair of electrons are counted with terminal S-atom.

I, S-atom : Number of electrons in the valence shell = 6

Number of electrons left after bonding = 0

Oxidation number of central S-atom = 6 – 0 = + 6

II, S-atom : Number of electrons in the valence shell = 6

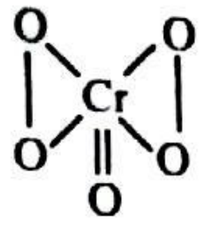
Number of electrons left after bonding = 8

Oxidation number of terminal S-atom = 6 – 8 = – 2

Now, you can also calculate Average Oxidation number of $S = \frac{6+(-2)}{2} = +2$

Special Examples of oxidation state

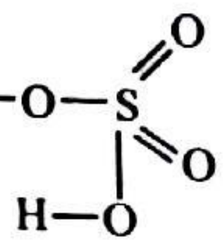
In order to determine the exact or individual oxidation number we need to take help from the structures of the molecules. Some special cases are discussed as follows:

1. The structure of CrO_5 is 

From the structure, it is evident that in CrO_5 there are two peroxide linkages and one double bond. The contribution of each peroxide linkage is –1. Let the oxidation number of Cr is x.

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

$\therefore$ Oxidation number of Cr = + 6 **Ans**

2. The structure of H_2SO_5 is 

From the structure, it is evident that in H_2SO_5, there is one peroxide linkage, two sulphur-oxygen double bonds and one OH group. Let the oxidation number of S = x.

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

or $x + 2 - 8 = 0$ or $x - 6 = 0$

or $x = 6$

$\therefore$ Oxidation number of S in H_2SO_5 is + 6

Paradox of Fractional Oxidation Number

Fractional oxidation number is the average of oxidation state of all atoms of element under examination and the structural parameters reveal that the atoms of element for whom fractional oxidation state is determined is actually present in different oxidation states.

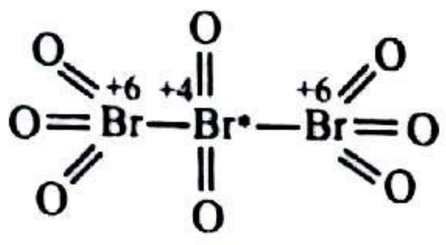
Structure of the species C_3O_2, Br_3O_8 and $S_4O_6^{2-}$ reveal the following bonding situations:

The element marked with asterisk (*) in each species is exhibiting different oxidation number from rest of the atoms of the same element in each of the species. This reveals that in C_3O_2, two carbon atoms are present in +2 oxidation state each whereas the third one is present in zero oxidation state and the average is + 4/3. However, the realistic picture is +2 for two terminal carbons and zero for the middle carbon.

$$O = \overset{+2}{C} = \overset{0}{C}{}^{*} = \overset{+2}{C} = O$$

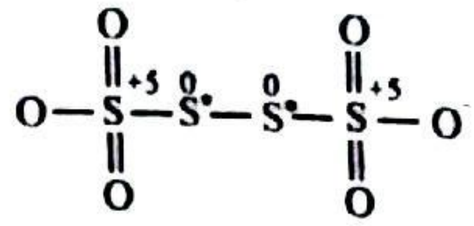
Structure of C_3O_2
(Carbon suboxide)

Likewise in Br_3O_8, each of the two terminal bromine atoms are present in +6 oxidation state and the middle bromine is present in +4 oxidation state. Once again the average, that is different from reality, is + 16/3.



Structure of Br_3O_6 (Tribromooctaoxide)

In the species $S_4O_6^{2-}$, average oxidation number of S is +2.5, whereas the reality being +5,0,0 and +5 oxidation number respectively for respective sulphur atoms.



Structure of $S_4O_6^{2-}$ (tetrathionate ion)

In general, the conclusion is that the idea of fractional oxidation state should be taken with care and the reality is revealed by the structures only.

Applications of Oxidation Number

To compare the strength of acid and base:

Strength of acid $\propto$ Oxidation Number

Strength of base $\propto \frac{1}{\text{Oxidation Number}}$

Example: Order of acidic strength in HClO, $HClO_2$, $HClO_3$, $HClO_4$ will be.

	Oxidation Number of chlorine
HClO (Hypo chlorous acid)	+1
$HClO_2$ (Chlorous acid)	+3
$HClO_3$ (Chloric acid)	+5
$HClO_4$ (Perchloric acid)	+7

Strength of acid α Oxidation Number

So the order will be

$$HClO_4 > HClO_3 > HClO_2 > HClO$$

To determine the oxidising and reducing nature of the substances:

(*a*) If effective element in a compound is present in maximum Oxidation state then the compound acts as oxidising agent.

Ex.

$KMnO_4$	$K_2Cr_2O_7$	H_2SO_4	SO_3	H_3PO_4	HNO_3	$HClO_4$
↓	↓	↓	↓	↓	↓	↓
+7	+6	+6	+6	+5	+5	+7

(*b*) If effective element in a compound is present in intermediate Oxidation state then the compound can act as oxidising agent as well as reducing agent.

HNO_2	H_3PO_3	SO_2	H_2O_2
↓	↓	↓	↓
+3	+3	+4	–1

To determine the molecular formula of the compound.

To calculate the equivalent weight of compounds:

The equivalent weight of an oxidising agent is that weight which accepts one mole electrons in a chemical reaction.

(*a*) Equivalent weight of an oxidant (get reduced)

$$= \frac{\text{Molecular weight}}{\text{No. of electrons gained by one mole}}$$

Example: In acidic medium

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

Here atoms which undergoes reduction is Cr. Its O.S. is decreasing from +6 to +3

$$\text{Eq. wt. of } K_2Cr_2O_7 = \frac{\text{Mol. wt of } K_2Cr_2O_7}{3 \times 2} = \frac{\text{Mol.wt.}}{6}$$

6 in denominator indicates that 6 electrons were gained by $Cr_2O_7^{2-}$ as it is clear from the given balanced eq.

(b) Similarly equivalent wt. of a reductant (gets oxidised)

$$\frac{\text{Mol. wt.}}{\text{No. of electrons lost by one mole}}$$

In acidic medium, $C_2O_4^{2-} \rightarrow 2CO_2 + 2e$

Here, Total electrons lost = 2

So, eq wt. $= \frac{\text{Mol. wt.}}{2}$

(c) In different conditions a compound may have different equivalent weights. Because, it depends upon the number of electrons gained or lost by that compound in that reaction.

Acid-base reaction

In case of acid base reaction, the valence factor is the actual number of H^+ or OH^- replaced in the reaction. The acid or base may contain more number of replaceable H^+ or OH^- than actually replaced in reaction.

Valency factor for base is the number of H^+ ion from the acid replaced by each molecule of the base.

KEY NOTE

Valency factor calculation:

- **For Elements:** Valency factor = valency of the element.
- **For Acids:** Valency factor = number of replaceable H^+ ions per acid molecule

TRAIN YOUR BRAIN

Q. $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
Base Acid

Ans. Valency factor of base = 1

Here, two molecule of NaOH replaced $2H^+$ ion from the H_2SO_4. Therefore, each molecule of NaOH replaced only one H^+ ion of acid, so valency factor = 1.

Salts:

(1) In non-reacting condition

Valency factor = Total number of positive charge or negative charge present in the compound.

TRAIN YOUR BRAIN

Q. Calculate valency factor for the following

	Na_2CO_3,	$Fe_2(SO_4)_3$	$FeSO_4.7H_2O$
Ans. V.f.	2	$2 \times 3 = 6$	2
Eq.wt.	$\frac{M}{2}$	$\frac{M}{6}$	$\frac{M}{2}$

(2) In reacting condition

TRAIN YOUR BRAIN

Q. $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$
Base Acid

Ans. It is an acid base reaction, therefore valency factor for Na_2CO_3 is one while in non-reacting condition, it will be two.

(3) Equivalent weight of oxidising/reducing agents in a redox reaction

In case of redox change,

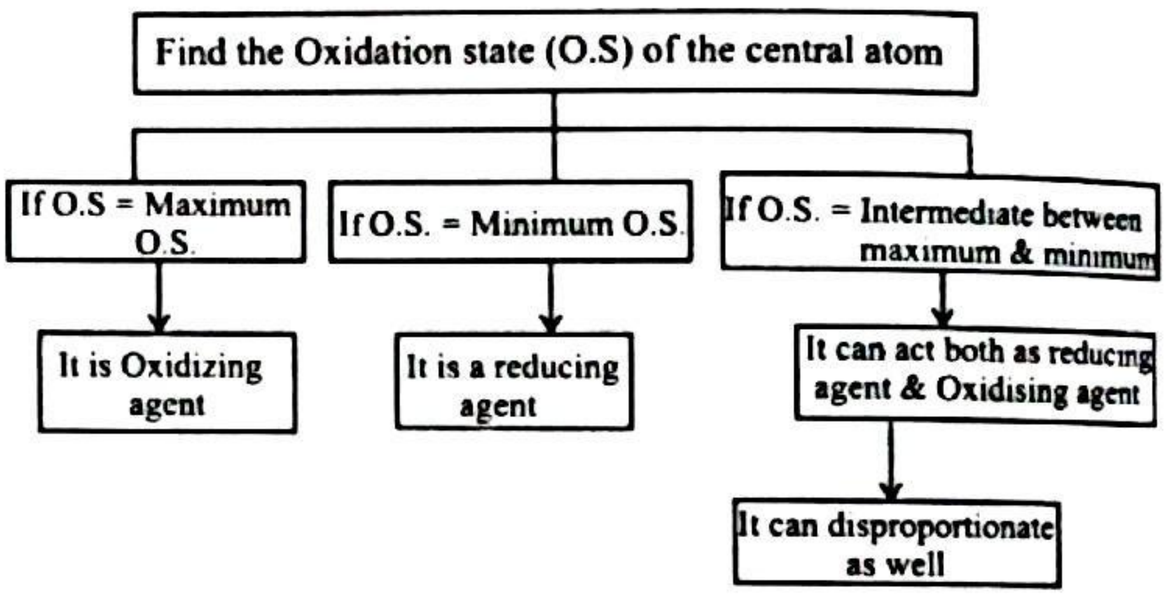
v.f. = Total change in oxidation number per molecule.

TRAIN YOUR BRAIN

Q. $KMnO_4 + H_2O_2 \rightarrow Mn^{2+} + O_2$

Ans. Mn in $KMnO_4$ is going from +7 to +2, so change in oxidation number per molecule of $KMnO_4$ is 5. So the valency factor of $KMnO_4$ is 5 and equivalent weight is $\frac{M}{5}$.

How To Identify Whether A Particular Substance Is An Oxidising Or A Reducing Agent



Redox reaction: A reaction in which oxidation and reduction simultaneously take place is called a redox reaction

In all redox reactions, the total increase in oxidation number must be equal to the total decrease in oxidation number.

$$10\overset{+2}{Fe}SO_4 + 2K\overset{+5}{Mn}O_4 + 8H_2SO_4 \rightarrow 5\overset{+3}{Fe_2}(SO_4)_3 + 2\overset{+2}{Mn}SO_4 + K_2SO_4 + 8H_2O$$

TYPES OF REDOX REACTIONS

1. **Chemical combination reactions:** The redox reactions in which the compounds are formed by combining two or more elements with each other or with a compound.

   $A + B \rightarrow AB$

   Examples:

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

   $$\overset{0}{S}(s) + \overset{0}{O_2}(g) \rightarrow \overset{+4\ -2}{S\,O_2}(g)$$

   $$\overset{-4\ +1}{C\,H_4}(g) + \overset{0}{O_2}(g) \rightarrow CO_2(g) + 2\overset{+1\ -2}{H_2\,O}(l)$$

2. **Decomposition reactions:** The redox reactions in which the chemical compounds chemically split into two or more simple substances.

   $AB \rightarrow A + B$

   Example: $2\overset{+1\ +5}{K\,Cl\,O_3^{-2}} \rightarrow 2\overset{+1\ -1}{K\,Cl} + 3\overset{0}{O_2}$

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

3. **Displacement reactions**

(*a*) The redox reactions in which the place of one species in compound is taken up by other species

E.g: $A + BC \rightarrow AC + B$

$$\overset{0}{Zn}(s) + Cu^{+2}SO_4^{-2}(aq) \rightarrow Zn^{+2}SO_4^{-2}(aq) + \overset{0}{Cu}(s)$$

(b) In non metal displacement redox reactions, generally hydrogen gets displaced and rarely oxygen.

E.g: $\overset{0}{Na}(s) + 2\overset{+1}{H}_2\overset{-2}{O}(l) \rightarrow 2\overset{+1}{Na}\overset{-2+1}{OH}(aq) + \overset{0}{H}_2(g)$

$$\overset{0}{Zn}(s) + 2\overset{+1}{H}\overset{-1}{Cl}(aq) \rightarrow \overset{+2}{Zn}\overset{-1}{Cl}_2(aq) + \overset{0}{H}_2(g)$$

4. **Disproportionation Reaction:** A redox reaction in which same element present in a particular compound in a definite oxidation state is oxidized as well as reduced simultaneously is a disproportionation reaction.

Disproportionation reactions are a special type of redox reactions. One of the reactants in a disproportionation reaction always contains **an element that can exist in at least three oxidation states**. The element in the form of reacting substance is in the intermediate oxidation state and both higher and lower oxidation states of that element are formed in the reaction. For example:

$$2\overset{-1}{H_2O_2}(aq) \rightarrow 2\overset{-2}{H_2O}(l) + \overset{0}{O_2}(g)$$

$$\overset{0}{S_8}(s) + 12OH^-(aq) \rightarrow 4\overset{-2}{S^{2-}}(aq) + 2\overset{+2}{S_2}O_3^{2-}(aq) + 6H_2O(\ell)$$

✍ KEY NOTE

Reverse of disproportionation is called **Comproportionation.** In some of the disproportionation reactions, by changing the medium (from acidic to basic or reverse), the reaction goes in backward direction and can be taken as an example of **Comproportionation reaction.**

$$\overset{+5}{HIO_3} + 5\overset{-1}{HI} \rightarrow 3\overset{0}{I_2} + 3H_2O$$

Consider the following reactions:

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

$KClO_3$ plays a role of oxidant and reductant both. Here, Cl present in $KClO_3$ is reduced and O present in $KClO_3$ is oxidized. Since same element is not oxidized and reduced, so **it is not a disproportionation reaction,** although it looks like one.

List of some important disproportionation reactions

- $H_2O_2 \rightarrow H_2O + O_2$
- $X_2 + OH^-(dil.) \rightarrow X^- + XO^-$ (X = Cl, Br, I)
- $X_2 + OH^-(conc.) \rightarrow X^- + XO_3^-$
- $(CN)_2 + OH^- \rightarrow CN^- + OCN^-$
- $P_4 + OH^- \rightarrow PH_3 + H_2PO_2^-$
- $S_8 + OH^- \rightarrow S^{2-} + S_2O_3^{2-}$
- $MnO_4^{2-} \rightarrow MnO_4^- + MnO_2$
- $NH_2OH \rightarrow N_2O + NH_3$
- $NH_2OH \rightarrow N_2 + NH_3$
- Oxyacids of Phosphorus (+1, +3 oxidation number)
- $H_3PO_2 \rightarrow PH_3 + H_3PO_3$
- $H_3PO_3 \rightarrow PH_3 + H_3PO_4$
- Oxyacids of Chlorine (Halogens) (+1, +3, +5 Oxidation number)
- $ClO^- \rightarrow Cl^- + ClO_2^-$
- $ClO_2^- \rightarrow Cl^- + ClO_3^-$
- $ClO_3^- \rightarrow Cl^- + ClO_4^-$
- $HNO_2 \rightarrow NO + HNO_3$

✍ KEY NOTE

F_2 does not undergo disproportionation as it is the most electronegative element and does not show positive oxidation state.

BALANCING OF REDOX REACTION

There are two ways to balance redox reaction:-

(1) Oxidation number change method.

(2) Ion electron method. (Half reaction method)

(1) Oxidation number change method: This method was given by Johnson. In a balanced redox reaction, total increase in oxidation number must be equal to total decreases in oxidation number. This equivalence provides the basis for balancing redox reactions.

The general method involves the following steps:

(*i*) Select the atom in oxidising agent whose oxidation number decreases and indicate the gain of electrons.

(*ii*) Select the atom in reducing agent whose oxidation number increases and indicate the loss of electrons.

(*iii*) Now cross multiply i.e., multiply oxidising agent by the number of loss of electrons and reducing agent by number of gain of electrons.

(*iv*) Balance the number of atoms on both sides whose oxidation numbers changes in the reaction.

(*v*) In order to balance oxygen atoms, add H_2O molecules to the side deficient in oxygen.

(*vi*) Then balance the number of H atoms by adding H^+ ions to the side deficient in hydrogen.

(*vii*) This is the balanced equation in acidic medium for basic medium, add ^-OH ion same in no with H^+ and simplify the equation.

TRAIN YOUR BRAIN

Q. Balance the following reaction by the oxidation number method–

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

Ans. Write the oxidation number of all the atoms.

$$\overset{0}{Cu} + \overset{+1+5-2}{HNO_3} \rightarrow \overset{+2+5-2}{Cu(NO_3)_2} + \overset{+4-2}{NO_2} + \overset{+1-2}{H_2O}$$

There is change in oxidation number of Cu and N.

$$\overset{0}{Cu} \rightarrow \overset{+1+5+2}{Cu(NO_3)_2} \text{ ...(1) (Oxidation no. is increased by 2)}$$

$$\overset{+5}{HNO_3} \rightarrow \overset{+4}{NO_2} \text{ ...(2) (Oxidation no. is decreased by 1)}$$

To make increase and decrease equal, eq. (2) is multiplied by 2.

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

Balancing nitrates ions, hydrogen and oxygen, the following equation is obtained.

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

This is the balanced equation.

(2) Ion-Electron method: This method was given by Jette and La mev in 1972.

The following steps are followed while balancing redox reaction (equations) by this method.

(*i*) Write the equation in ionic form.

(*ii*) Split the redox equation into two half reactions, one representing oxidation and the other representing reduction.

(*iii*) Balance these half reactions separately and then add by multiplying with suitable coefficients so that the electrons are cancelled. Balancing is done using following substeps:

(*a*) Balance all other atoms except H and O.

(*b*) Then balance oxygen atoms by adding H_2O molecules to the side deficient in oxygen. The number of H_2O molecules added is equal to the deficiency of oxygen atoms.

(*c*) Balance hydrogen atoms by adding H^+ ions equal to the deficiency in the side which is deficient in hydrogen atoms.

(*d*) Balance the charge by adding electrons to the side which is rich in +ve charge. i.e. deficient in electrons. Number of electrons added is equal to the deficiency.

(*e*) Multiply the half equations with suitable coefficients to equalize the number of electrons.

(*iv*) Add these half equations to get an equation which is balanced with respect to charge and atoms.

(*v*) If the medium of reaction is basic, OH^- ions are added to both sides of balanced equation, which is equal to number of H^+ ions in Balanced Equation.

Example:

(i) Acidic Medium:

(*a*) $Cr_2O_7^{2-} + C_2O_4^{2-} \xrightarrow{H^+} Cr^{3+} + CO_2$

(*b*) Write both the half reaction.

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (Reduction half reaction)

$C_2O_4^{2-} \rightarrow CO_2$ (Oxidation half reaction)

(*c*) Atoms other than H and O are balanced.

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

$C_2O_4^{2-} \rightarrow 2CO_2$

(*d*) Balance O-atoms by the addition of H_2O to another side

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

$C_2O_4^{2-} \rightarrow 2CO_2$

(*e*) Balance H-atoms by the addition of H^+ to another side

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

$C_2O_4^{2-} \rightarrow 2CO_2$

(*f*) Now, balance the charge by the addition of electron (e^-).

$Cr_2O_7^{2-} + 14\,H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ .....(1)

$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$ .....(2)

(*g*) Multiply equations by a constant to get the same number of electrons on both side. In the above case second equation is multiplied by 3 and then added to first equation.

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

$3C_2O_4^{2-} \rightarrow 6CO_2 + 6e$

$Cr_2O_7^{2-} + 3C_2O_4^{2-} + 14\,H^+ \rightarrow 2Cr^{3+} + 6CO_2 + 7H_2O$

(ii) Alkaline Medium:

(*a*) Consider the reaction

$Cr(OH)_3 + IO_3^- \xrightarrow{OH} I^- + CrO_4^{2-}$

(*b*) Separate the two half reactions.

$Cr(OH)_3 \rightarrow CrO_4^{2-}$ (Oxidation half reaction)

$IO_3^- \rightarrow I^-$ (Reduction half reaction)

(*c*) Balance O-atoms by adding H_2O.

$H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-}$

$IO_3^- \rightarrow I^- + 3H_2O$

(*d*) Balance H-atoms by adding H^+ to side having deficiency and add equal no. of OH^- ions to both the sides (medium is alkaline)

$H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+$

$5OH^- + H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 5OH^-$

or $5OH^- + Cr(OH)_3 \rightarrow CrO_4^{2-} + 4H_2O$

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

or $IO_3^- + 3H_2O \rightarrow I^- + 6OH^-$

(*e*) Balance the charges by adding electrons

$5OH^- + Cr(OH)_3 \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$

$IO_3^- + 6H_2O + 6e^- \rightarrow I^- + 3H_2O + 6OH^-$

(*f*) Multiply first equation by 2 and add to second to give

$10OH^- + 2Cr(OH)_3 \rightarrow 2\,CrO_4^{2-} + 8H_2O + 6e^-$

$IO_3^- + 6H_2O + 6e^- \rightarrow I^- + 3H_2O + 6OH^-$

Add there balanced half reaction to give complete balanced reaction.

$4OH^- + 2Cr(OH)_3 + IO_3^- \rightarrow 5H_2O + 2Cr_2O_4^{2-} + I^-$

APPLICATIONS OF REDOX REACTIONS IN QUANTITATIVE ANALYSIS

In titrimetric quantitative analysis:

The solutions of accurately known concentration is called standard solution.

In titrimetric analysis, the substance of known concentration is called titrant.

Substanction is detected by

(1) Observing physical change

(2) By using an Indicator.

The usage of indicators in redox titration is shown below:

- In $Cr_2O_7^{2-}$ (dichromate) titrations, diphenyl amine is used as a reagent. End point is intense blue colour due to reduction of $Cr_2O_7^{2-}$.
- In titration of Cu^{2+} and I^- to I_2 it has deep blue Colour with starch solution due to oxidation of I^-

$$Cu^{2+}(aq) + 4I^-(aq) \rightarrow Cu_2I_2(s) + I_2(aq)$$

- In the titration of I_2 with thiosulphate, blue colour disappears due to the conversion of I_2 to I^-.

Table: Some Redox titratrion Excluding Iodometric/ Iodimetric

	Estimation of	By titrating with	Reaction	Relation* between OA and RA
1.	Fe^{2+}	MnO_4^-	$Fe^{2+} \rightarrow Fe^{3+} + e^-$ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$5Fe^{2+} \equiv MnO_4^-$ Eq. wt. of Fe^{2+} = M/1
2.	Fe^{2+}	$Cr_2O_7^{2-}$	$Fe^{2+} \rightarrow Fe^{3+} + e^-$ $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	$6Fe^{2+} \equiv Cr_2O_7^{2-}$ Eq. wt. of $Cr_2O_7^{2-}$ = M/6
3.	$C_2O_4^{2-}$	MnO_4^-	$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$5C_2O_4^{2-} \equiv 2MnO_4^-$ Eq. wt. of $C_2O_4^{2-}$ = M/2
4.	H_2O_2	MnO_4^-	$H_2O_2 \rightarrow 2H^+ + O_2 + 2e^-$ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$5H_2O_2 \equiv 2MnO_4^-$ Eq. wt. of H_2O_2 = M/2
5.	As_2O_3	MnO_4^-	$As_2O_3 + 5H_2O \rightarrow 2AsO_4^{3-} + 10H^+ + 4e^-$ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	Eq. wt. of As_2O_3 = M/4
6.	AsO_3^{3-}	BrO_3^-	$AsO_3^{3-} + H_2O \rightarrow AsO_4^{3-} + 2H^+ + 2e^-$ $BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O$	Eq. wt of AsO_3^{3-} = M/2 Eq. wt. of BrO_3^- = M/6

TRAIN YOUR BRAIN

Q. Write the balanced reaction of titration of $KMnO_4$ Vs oxalic acid in presence of H_2SO_4.

Ans. Reaction : $2KMnO_4 + 3H_2SO_4 + 5H_2C_2O_4 \rightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 10CO_2$

Redox Changes :

$$C_2^{3+} \rightarrow 2C^{4+} + 2e^- \quad \left(E_{H_2C_2O_4} = \frac{M}{2}\right)$$

$$5e^- + Mn^{7+} \rightarrow Mn^{2+} \quad \left(E_{KMnO_4} = \frac{M}{5}\right)$$

Indicator : $KMnO_4$ acts as self indicator.

REDOX REACTION AND ELECTRODE PROCESS

Electrode Potential

- The potential difference developed between metal electrode and its ions in solution is known as **electrode potential**.

Electrode Potential

Oxidation Potential (O.P.):
- The electrode potential for oxidation half reaction
- Tendency to get oxidised
- Greater the O.P. then greater will be tendency to get oxidised.

Reduction Potential(R.P.)
- The electrode potential for reduction half reaction
- Tendency to get reduced.
- Greater the R.P., greater will be tendency to get reduced.

Standard Electrode Potential (E°)

The potential difference developed between metal electrodes and the solution of its ions at 1 M concentration at 1 bar pressure and 298 K is **known as standard electrode potential** or standard **reduction potential (old name).**

KEY NOTE

- According to IUPAC, the given value of electrode potential is regarded as reduction potential unless it is specifically mentioned.
- The **reduction potential** of an electrode is exactly equal in magnitude but opposite in sign to its **oxidation potential.**

Electrochemical Series

- In the electrochemical series, the metals are arranged in the **increasing order of reduction potentials.**
- High reduction potential, metal ions undergo reduction first on cathode or it acts as cathode in the electrochemical cell.
- Low reduction potential, metal (high negative value) undergo oxidation or it acts as anode. It is more reactive or more electropositive.
- In the electrochemical series **metal placed above** can displace the **below metal ions** from their solution.

Table: Electrochemical Series The standard Electrode potantial

Electrode	Reaction	SRP (at 298 K)
*Li	$Li^+ + e^- \rightarrow Li(s)$	–3.05 V
K	$K^+ + e^- \rightarrow K$	–2.93 V
Ca	$Ca^{+2} + 2e^- \rightarrow Ca(s)$	–2.87 V
Na	$Na^+ + e^- \rightarrow Na(s)$	–2.71 V
Mg	$Mg^{+2} + 2e^- \rightarrow Mg(s)$	–2.37 V
Al	$Al^{3+} + 3e^- \rightarrow Al(s)$	–1.66 V
* Electrolytes (H_2O)	$H_2O(l) + e^- \rightarrow \frac{1}{2}H_2 + OH^-$	–0.83 V
*Zn	$Zn^{+2} + 2e^- \rightarrow Zn(s)$	–0.76 V
Cr	$Cr^{+3} + 3e^- \rightarrow Cr(s)$	–0.74 V
* Fe	$Fe^{2+} + 2e^- \rightarrow Fe(s)$	–0.44 V
Cd	$Cd^{+2} + 2e^- \rightarrow Cd(s)$	–0.40 V
Co	$Co^{+2} + 2e^- \rightarrow Co(s)$	–0.277 V
Ni	$Ni^{+2} + 2e^- \rightarrow Ni(s)$	–0.24 V
Sn	$Sn^{+2} + 2e^- \rightarrow Sn(s)$	–0.14 V
Pb	$Pb^{+2} + 2e^- \rightarrow Pb(s)$	–0.13 V
* H_2	$2H^+ + 2e^- \rightarrow H_2(g)$	0.00 V
Cu	$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34 V
I_2	$I_2 + 2e^- \rightarrow 2I^-$	0.54 V
Fe	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77 V
Hg	$Hg_2^{2+} + 2e^- \rightarrow 2Hg(l)$	0.79 V
Ag	$Ag^+ + e^- \rightarrow Ag(\ell)$	0.80 V
Hg	$Hg^{2+} + 2e^- \rightarrow Hg(\ell)$	0.85 V
Br_2	$Br_2 + 2e^- \rightarrow 2Br^-$	1.06 V
* Electrolytes	$\frac{1}{2}O_2 + 2H^+ + 2e^- \rightarrow H_2O(\ell)$	1.23 V
* $Cr_2O_7^{2-}$	$+ 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$	1.33 V
* Cl_2	$+ 2e^- \rightarrow 2Cl^-$	1.36 V
* MnO_4^-	$+ 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51 V
* $F_2(g)$	$+ 2e^- \rightarrow 2F^-$	2.87 V

Increasing strength of reducing agent

Increasing strength of oxidising agent

Topicwise Questions

CLASSICAL IDEA OF REDOX REACTIONS

1. In oxidation process:
   (*a*) Oxidation number decreases
   (*b*) Number of electron decrease
   (*c*) Oxygen content decreases
   (*d*) Number of ions decreases

2. Which of the following is the most powerful oxidising agent?
   (*a*) F_2 (*b*) O_2
   (*c*) Br_2 (*d*) I_2

3. In which of the following reactions, H_2O_2 acts as a reducing agent?
   (*a*) $HOCl + H_2O_2 \xrightarrow{H^+} H_3O^+ + Cl^- + O_2$
   (*b*) $Mn^{2+} + H_2O_2 \xrightarrow{OH^-} Mn^{4+} + 2OH^-$
   (*c*) $2Fe^{2+} + H_2O_2 \xrightarrow{OH} 2Fe^{3+} + 2OH^-$
   (*d*) $PbS(s) + 4H_2O_2(aq) \xrightarrow{H^+} PbSO_4(s) + 4H_2O(l)$

4. In the reaction,
   $Ag_2O + H_2O_2 \rightarrow 2Ag + H_2O + O_2$
   H_2O_2 ACT AS:
   (*a*) Reducing agent (*b*) Oxidising agent
   (*c*) Bleaching agent (*d*) None of these

5. In $Cu^{2+} + Ag \rightarrow Cu + Ag^+$, oxidation half-reaction is:
   (*a*) $Cu^{2+} \rightarrow Cu$ (*b*) $Ag \rightarrow Ag^+$
   (*c*) $Cu \rightarrow Cu^{2+}$ (*d*) All of these

6. In one of the following reactions, HNO_3 does not behave as an oxidising agent, identify it:
   (*a*) $I_2 + 10\,HNO_3 \rightarrow 2\,HIO_3 + 10NO_2 + 4H_2O$
   (*b*) $3Cu + 8\,HNO_3 \rightarrow 3Cu\,(NO_3)_2 + 2NO + 4H_2O$
   (*c*) $4Zn + 10HNO_3 \rightarrow 4Zn\,(NO_3)_2 + NH_4NO_3 + 3H_2O$
   (*d*) $2HNO_3 + P_2O_5 \rightarrow 2HPO_3 + N_2O_5$

7. The equivalent weight of $MnSO_4$ is equal to its molecular weight when it is converted to:
   (*a*) Mn_2O_3 (*b*) MnO_2
   (*c*) MnO_4^- (*d*) MnO_4^{2-}

8. In which of the following reactions, the underlined substance has been oxidised?
   (*a*) $\underline{Br_2} + H_2S \rightarrow 2HBr + S$
   (*b*) $\underline{2HgCl_2} + SnCl_2 \rightarrow Hg_2Cl_2 + SnCl_4$
   (*c*) $Cl_2 + \underline{2KI} \rightarrow 2KCl + I_2$
   (*d*) $\underline{2Cu^{2+}} + 4I^- \rightarrow Cu_2I_2 + I_2$

9. Which changes requires an oxidising agent?
   (*a*) $2S_2O_3^{2-} \rightleftharpoons S_4O_6^{2-}$ (*b*) $Zn^{2+} \rightleftharpoons Zn$
   (*c*) $ClO^- \rightleftharpoons Cl^-$ (*d*) $SO_3 \rightleftharpoons SO_4^{2-}$

10. Which of the following molecules can act as an oxidising as well as reducing agent?
    (*a*) H_2S (*b*) SO_3
    (*c*) H_2O_2 (*d*) F_2

11. Which of the following is a redox reaction?
    (*a*) $2CuSO_4 + 4KI \rightarrow Cu_2I_2 + 2K_2SO_4 + I_2$
    (*b*) $SO_2 + H_2O \rightarrow H_2SO_3$
    (*c*) $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$
    (*d*) $CuSO_4 + 4NH_3 \rightarrow [Cu(NH_3)_4]\,SO_4$

12. According to classical concept, oxidation involves
    (*a*) Addition of oxygen
    (*b*) Addition of electronegative element
    (*c*) Removal of either hydrogen or some electropositive element
    (*d*) All of these

REDOX REACTIONS IN TERMS OF ELECTRON TRANSFER REACTIONS

13. What volume of 0.1 N oxalic acid solution can be reduced by 250 g of an 8 percent by weight $KMnO_4$ solution?
    (*a*) 6.3 litre (*b*) 12.6 litre
    (*c*) 25.2 litre (*d*) 0.63 litre

14. In the reaction,
    $N_2 \rightarrow NH_3$.
    The eq. wt. of N_2 and NH_3 are respectively equal to:
    (*a*) $\frac{28}{3}, \frac{17}{3}$ (*b*) $\frac{28}{6}, \frac{17}{3}$
    (*c*) $\frac{28}{2}, \frac{17}{2}$ (*d*) $\frac{28}{5}, \frac{17}{5}$

15. A 0.518 g sample of lime stone is dissolved in HCl and then the calcium is precipitated as CaC_2O_4. After filtering and washing the precipitate, it requires 40.0 mL of 0.250 N $KMnO_4$, solution acidified with H_2SO_4 to titrate it as,
    $MnO_4^- + H^+ + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2 + 2H_2O$
    The percentage of CaO in the sample is:
    (*a*) 54.0% (*b*) 27.1%
    (*c*) 42% (*d*) 84%

16. A 0.50 M solution of KI reacts with excess of H_2SO_4 in KIO_3 solutions according to the equation,
    $6H^+ + 5I^- + IO_3^- \rightarrow 3I_2 + 3H_2O$.

Which of the following statements is true?

(*a*) 200 mL of the KI solution reacts with 0.10 mole KIO_3

(*b*) 100 mL of the KI solution reacts with 0.060 M of H_2SO_4

(*c*) 0.5 litre of the KI solutions produces 0.15 mole of I_2

(*d*) None of the above

17. In the following change,

$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

If the atomic weight of iron is 56, then its equivalent weight will be:

(*a*) 42 (*b*) 21

(*c*) 63 (*d*) 84

OXIDATION NUMBER

18. The oxidation number and covalency of sulphur in sulphur molecules (S_8) are:

(*a*) 0 and 2 (*b*) + 6 and 8

(*c*) 0 and 8 (*d*) + 6 and 2

19. Oxidation number of N in HNO_3 is:

(*a*) –3.5 (*b*) +3.5

(*c*) –5 (*d*) +5

20. In which of the following compounds, iron has an oxidation number of +3?

(*a*) $Fe(NO_3)_2$ (*b*) FeC_2O_4

(*c*) $FeCl_3$ (*d*) None of these

21. Oxidation states of P in $H_4P_2O_5$, $H_4P_2O_6$ $H_4P_2O_7$, respectively are:

(*a*) + 3, + 4, + 5 (*b*) + 3, + 5, + 4

(*c*) + 5, + 3, + 4 (*d*) + 5, + 4, + 3

22. 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

23. The oxidation number of cobalt in $K[Co(CO)_4]$ is:

(*a*) + 1 (*b*) +3

(*c*) –1 (*d*) –3

24. Oxidation number of iodine in IO_3^-, IO_4^-, KI and I_2 respectively are:

(*a*) –1, –1, 0 + 1 (*b*) + 3, + 5, + 7, 0

(*c*) + 5, + 7, –1, 0 (*d*) –1, –5, –1, 0

25. When sulphur dioxide is passed in an acidified $K_2Cr_2O_7$ solution, the oxidation state of sulphur is changed from:

(*a*) + 4 to 0 (*b*) + 4 to + 2

(*c*) + 4 to + 6 (*d*) + 6 to + 4

26. Of the following other electronic configurations of atoms, the highest oxidation state is achieved by which one of them?

(*a*) $(n-1)d^8\,ns^2$ (*b*) $(n-1)\,d^5\,ns^2$

(*c*) $(n-1)\,d^3\,ns^2$ (*d*) $(n-1)\,d^5\,ns^1$

27. What is the oxidation state of P in $Ba(H_2PO_2)_2$?

(*a*) +1 (*b*) +2

(*c*) +3 (*d*) –1

28. When a manganous salt is fused with the mixture of KNO_3 and the solid NaOH, the oxidation number of Mn changes from +2 to:

(*a*) +4 (*b*) +3

(*c*) +6 (*d*) +7

29. 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:

(*a*) +5 (*b*) +3

(*c*) +6 (*d*) –10

30. 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

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

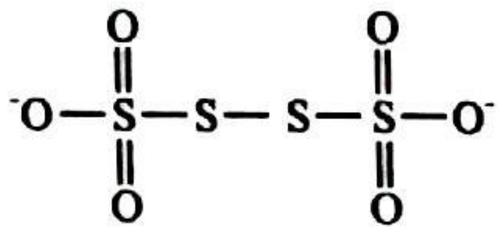
(*a*) $Na_2S_4O_6 > H_2S_2O_7 > Na_2S_2O_3 > S_8$

(*b*) $H_2SO_4 > SO_2 > H_2S > H_2S_2O_8$

(*c*) $SO_2 > SO_4^{2-} > SO_3^{2-} > HSO_4^-$

(*d*) $H_2SO_4 > H_2SO_3 > SCl_2 > H_2S$

32. The oxidation states of S-atoms in $S_4O_6^{2-}$ form left to right respectively are:



(*a*) + 6, 0, 0, + 6 (*b*) + 3, 1, + 1, + 3

(*c*) + 5, 0, 0, + 5 (*d*) +4, +1, + 1, + 4

33. In the conversion of Br_2 to BrO_3^-, the oxidation number of Br changes from:

(*a*) 0 to +5 (*b*) + 1 to +5

(*c*) 0 to – 3 (*d*) + 2 to + 5

34. The oxidation number of the sulphur atoms in peroxomonosulphuric acid H_2SO_5 and peroxodisulphuric ($H_2S_2O_8$) respectively, are:

(*a*) + 8 and + 7 (*b*) + 3 and + 3

(*c*) + 6 and + 6 (*d*) + 4 and + 6

35. The average oxidation number of Fe in Fe_3O_4 is:

(*a*) + 2 (*b*) +3

(*c*) $\frac{8}{3}$ (*d*) $\frac{2}{3}$

36. The reaction of aqueous $KMnO_4$ with H_2O_2 in acidic condition gives:

(*a*) Mn^{+4} and O_2 (*b*) Mn^{2+} and O_2

(*c*) Mn^{2+} and O_3 (*d*) Mn^{4+} and MnO_2

37. Carbon is in lowest oxidation state:

(*a*) CO_2 (*b*) CF_4

(*c*) CCl_4 (*d*) CH_4

38. Oxidation state of phosphorus in pyrophosphoric acid is:

(*a*) + 5 (*b*) + 3

(*c*) +4 (*d*) + 1

39. Oxidation state of nitrogen is correctly given for:

Compound	Oxidation state
(a) Mg_3N_2	– 3
(b) NH_2OH	+ 1
(c) $(N_2H_5)_2SO_4$	+2
(d) $[Co(NH_3)_5Cl]Cl_2$	0

40. Oxidation number of 'N' in N_3H (hydrazoic acid) is:

(a) –1/3 (b) + 3

(c) 0 (d) –3

41. Oxidation state of hydrogen in CaH_2 is-

(a) +1 (b) –1

(c) + 2 (d) 0

42. The oxidation number of P is + 3 in-

(a) H_3PO_3 (b) H_3PO_4

(c) HPO_3 (d) $H_4P_2O_7$

43. In the reaction,

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

(a) Bromine is oxidised and carbonate is reduced

(b) Bromine is oxidised as well as reduced

(c) Bromine is reduced and water is oxidised

(d) Br_2 is neither oxidised nor reduced

44. Which one of the following is not a redox reaction?

(a) $CaCO_3 \rightarrow CaO + CO_2$

(b) $2H_2 + O_2 \rightarrow 2H_2O$

(c) $Na + H_2O \rightarrow NaOH + \frac{1}{2}H_2$

(d) $MnCl_3 \rightarrow MnCl_2 + \frac{1}{2}Cl_2$

45. Which of following is not a redox change?

(a) $2H_2S + SO_2 \rightarrow 2H_2O + 3S$

(b) $2BaO + O_2 \rightarrow 2BaO_2$

(c) $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$

(d) $2KClO_3 \rightarrow 2KCl + 3O_2$

46. Which of the following halogens always shows only one oxidation state in compounds?

(a) Cl (b) F

(c) Br (d) I

47. The compound in which oxidation state of metal is zero-

(a) $Fe_2(CO)_9$ (b) Ni $(CO)_4$

(c) $Fe_3(CO)_9$ (d) All the above

48. In a reaction,

$H_2O + C \rightarrow CO + H_2$

(a) H_2O is the reducing agent

(b) H_2O is the oxidising agent

(c) Carbon is the oxidising agent

(d) Oxidation-reduction does not occur

49. The oxidation number of chlorine in HOCl is-

(a) – 1 (b) 0

(c) + 1 (d) 2

50. Which one can act as oxidising & reducing agent both?

(a) HNO_2 (b) H_2O_2

(c) H_2SO_3 (d) All

51. Which compound can not be used as oxidising agent?

(a) O_3 (b) HNO_3

(c) $KMnO_4$ (d) NH_3

52. Oxidation number of S in H_2SO_5 is 6. This is observed because-

(a) There are five oxygen atoms in the molecule

(b) The hydrogen atom is directly linked with non-metal

(c) There is peroxide linkage in the molecule

(d) The sulphur atom shows co-ordinate linkage

53. In which of the following compound oxidation number of Cl is + 3?

(a) ICl (b) ClO_3^-

(c) ClF_3 (d) $HClO_4$

REDOX REACTIONS AND ELECTRODE PROCESSES

54. In the redox reaction,

$xKMnO_4 + NH_3 \rightarrow y\,KNO_3 + MnO_2 + KOH + H_2O$

x and y are:

(a) x = 4, y = 6 (b) x = 3, y = 8

(c) x = 8, y = 6 (d) x = 8, y = 3

55. The number of electrons involved in the reduction of one nitrate ion to hydrazine is:

(a) 8 (b) 5

(c) 3 (d) 7

56. In ionic equation,

$BiO_3^- + 6H^+ + Xe^- \rightarrow Bi^{3+} + 3H_2O$

the value of x is:

(a) 6 (b) 2

(c) 4 (d) 3

57. In the following redox reaction,

$xUO^{2+} + Cr_2O_7^{2-} + yH^+ \rightarrow xUO_2^{2-} + zCr^{3+} + 6H_2O$

the value of coefficients x, y and z respectively are:

(a) 3, 8, 2 (b) 3, 8, 7

(c) 3, 2, 4 (d) 3, 1, 8

58. How many moles of iodine are liberated when 1 mole of potassium dichromate reacts with potassium iodide?

(a) 1 (b) 2

(c) 3 (d) 4

59. The equivalent mass of potassium permanganate in strong alkaline medium is:

(a) $\frac{\text{Molar mass}}{5}$ (b) $\frac{\text{Molar mass}}{3}$

(c) $\frac{\text{Molar mass}}{2}$ (d) Molar mass itself

60. In the reaction,

$IO_3^- + SO_2 + 4H_2O \rightarrow I_2 + SO_4^{2-} + 8H^+$

the coefficient of SO_2 is:

(*a*) 3 (*b*) 4

(*c*) 5 (*d*) 6

61. The coefficients of I^-, IO_3^- and H^+ in the redox reaction,

$I^- + IO_3^- + H^+ \rightarrow I_2 + H_2O$

In the balanced form respectively are:

(*a*) 5, 1, 6 (*b*) 1, 5, 6

(*c*) 6, 1, 5 (*d*) 5, 6, 1

62. In alkaline medium ClO_2 oxidises H_2O_2 to O_2 and reduced itself to Cl^-, then how many moles of H_2O_2 will oxidise by one mole of ClO_2?

(*a*) 1.0 (*b*) 1.5

(*c*) 2.5 (*d*) 3.5

63. For the redox reaction

$Zn + NO_3^- + H_2O \rightarrow Zn^{2+} + NH_4^+ + OH^-$

In the basic medium, coefficient of Zn, NO_3^- and OH^- in the balanced reaction respectively, are:

(*a*) 4, 1, 7

(*b*) 7, 4, 1

(*c*) 4, 1, 10

(*d*) 1, 4, 10

64. Which of the following examples does not represent disproportionation ?

(*a*) $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

(*b*) $2H_2O_2 \rightarrow 2H_2O + O_2$

(*c*) $4KClO_3 \rightarrow 3KClO_4 + KCl$

(*d*) $3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$

65. For the redox reaction

$MnO_4^- + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + H_2O$

In the balanced equation, correct coefficient are

	MnO_4^-	Fe^{2+}	H^+
(*a*)	1	5	8
(*b*)	16	5	2
(*c*)	5	16	2
(*d*)	2	16	5

66. Given : $Zn^{2+}/Zn = -0.76$ V, $Mg^{2+}/Mg = -2.37$ V, then the correct statement about the reaction $Zn(s) + MgCl_2(aq) \rightarrow$

(*a*) Solid zinc dissolves

(*b*) Zinc chloride precipitates

(*c*) Magnesium chloride precipitates

(*d*) No reaction takes place.

67. Equivalent weight of FeC_2O_4 in the change,

$FeC_2O_4 \rightarrow Fe^{3+} + CO_2$ is

(*a*) M/3 (*b*) M/6

(*c*) M/2 (*d*) M/1

Learning Plus

1. Which of the following is not an example of redox reaction?

(*a*) $CuO + H_2 \rightarrow Cu + H_2O$

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

(*c*) $2K + F_2 \rightarrow 2KF$

(*d*) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

2. The more positive the value of $E^\ominus$, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below find out which of the following is the strongest oxidizing agent?

$E^\ominus$ values : $Fe^{3+} / Fe^{2+} = +0.77$

$I_2(s)/I^- = +0.54$;

$Cu^{2+}/Cu = +0.34$; $Ag^+ / Ag = 0.80$ V

(*a*) Fe^{3+} (*b*) $I_2(s)$ (*c*) Cu^{2+} (*d*) Ag^+

3. $E^\ominus$ value of some redox couples are given below. On the basis of these values choose the correct option:

$E^\ominus$ values: $Br_2/Br^- = +1.90$

$Ag^+ /Ag(s) = +0.80$

$Cu^{2+}/Cu(s) = +0.34$; $I_2(s)/I^- = +0.54$

(*a*) Cu will reduce Br^- (*b*) Cu will reduce Ag

(*c*) Cu will reduce I^- (*d*) Cu will reduce Br_2

4. Using the standard electrode potential, find out the pair between which the action is not feasible $E^\ominus$ values:

$Fe^{3+}/Fe^{2+} = +0.77$; $I_2/I^- = +0.54$;

$Cu^{2+}/Cu = +0.34$; $Ag^+/Ag = +0.80$ V

(*a*) Fe^{3+} and I^- (*b*) Ag^+ and Cu

(*c*) Fe^{3+} and Cu (*d*) Ag and Fe^{3+}

5. This sulphate reacts differently with iodine and bromine in the reactions given below:

$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

$S_2O_3^{2-} + 2Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 4Br^- + 10H^+$

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*) This sulphate undergoes oxidation by bromine and reduction by iodine in these reactions

(*d*) Bromine undergoes oxidation and iodine undergoes reduction in these reactions

6. The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following is incorrect in this respect?

(*a*) The oxidation number of hydrogen is always +1

(*b*) The algebraic sum of all the oxidation numbers in a compound is zero

(*c*) An element in the free or the uncombined state bears oxidation number zero

(*d*) In all its compounds, the oxidation number of fluorine is –1

7. 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

8. Which of the following arrangements represent increasing oxidation number of the central atom?

(*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*) $CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-$

9. 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?

(*a*) $3d^1\ 4s^2$ (*b*) $3d^3\ 4s^2$

(*c*) $3d^5\ 4s^1$ (*d*) $3d^5\ 4s^2$

10. Identify disproportionation reaction:

(*a*) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(*b*) $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$

(*c*) $2F_2 + 2OH^- \rightarrow 2F^- + OF_2 + H_2O$

(*d*) $2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^- + H_2O$

11. Which of the following elements does not show disproportionation tendency?

(*a*) Cl (*b*) Br

(*c*) F (*d*) I

12. How many moles of MnO_4^- & $Ca(OH)_2$ are required to react with 1 mole of $Fe(HC_2O_4)_2$ respectively?

(*a*) 2, 1 (*b*) 3, 1 (*c*) 1, 1 (*d*) 2, 2

13. Equivalent weight of Br_2 is 96 in the following disproportionation reaction:

$Br_2 + OH^- \rightarrow Br^- + H_2O + X$

The oxidised product is: (Br = 80)

(*a*) BrO^- (*b*) BrO_2^-

(*c*) BrO_3^- (*d*) BrO_4^-

14. Calculate the equivalent weight of KHC_2O_4 when it is neutralised by KOH.

(*a*) 128 (*b*) 64

(*c*) 42 (*d*) None of these

15. Calculate the equivalent weight of KHC_2O_4 when it is oxidised by MnO_4^- in basic medium.

(*a*) 64 (*b*) 128

(*c*) 32 (*d*) 15

16. The number of mole of MnO_4^- required by 1 mole of ferric oxalate for oxidation is/are-

(*a*) 0.2 (*b*) 1.2

(*c*) 2.6 (*d*) 5.2

17. The oxidation states of S in SO_4^{2-}, $S_2O_8^{2-}$ & H_2SO_4 are respectively:

(*a*) +1, +6, +7 (*b*) +6, +7, +6

(*c*) +7, +7, +7 (*d*) +6, +6, +6

18. When $Cr(s) + OH^-(aq) \rightarrow Cr(OH)_4^-(aq) + H_2(g)$ (basic solution) is balanced, the sum of the coefficients of all the reactants and products is:

(*a*) 14 (*b*) 15

(*c*) 17 (*d*) 9

19. A 0.1097 g sample of As_2O_3 required 36.10 mL of $KMnO_4$ solution for its titration. The molarity of $KMnO_4$ solution is.

(*a*) 0.02 M (*b*) 0.04 M

(*c*) 0.0122 M (*d*) 0.3 M

20. 0.52 g of a dibasic acid required 100 mL of 0.2 N NaOH for complete neutralization. The equivalent weight of acid is

(*a*) 26 (*b*) 52

(*c*) 104 (*d*) 156

21. The minimum quantity of H_2S needed to precipitate 63.5 g of Cu^{2+} will be nearly.

(*a*) 63.5 g (*b*) 31.75 g

(*c*) 34 g (*d*) 2.0 g

22. Which of the following statements is/are true if 1 mol of H_3PO_x is completely neutralised by 40 g of NaOH?

(*a*) x = 2 and acid is monobasic

(*b*) x = 3 and acid is dibasic

(*c*) x = 3 and acid is tribasic

(*d*) x = 4 and given acid forms three series of salt

23. The weight of sodium bromate required to prepare 85.5 mL of 0.672 N solution for cell reaction,

$BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O$, is

(*a*) 1.56 g (*b*) 1.446 g

(*c*) 1.23 g (*d*) 1.32 g

24. If H_2S gas is passed through a solution of $K_2Cr_2O_7$, the colour of the solution will-

(*a*) Remain unchanged (*b*) Become deep red

(*c*) Become green (*d*) Become deep gray

25. How many electrons should X_2H_4 liberate so that in the new compound, X shows oxidation number of –1/2? [E.N. of X. > H]

(*a*) 10 (*b*) 4

(*c*) 3 (*d*) 2

26. The O.S. of oxygen in O_2PtF_6 is-

(*a*) Zero (*b*) + 1/2

(*c*) + 1 (*d*) - 1/2

27. Maximum oxidation state which can be shown by the elements of group 8 of the periodic table is

(*a*) +4 (*b*) +6

(*c*) +8 (*d*) 0

28. Equivalent weight of Mn^{3+} in the following reaction is (Mn = 55)

$Mn^{3+} \rightarrow Mn^{2+} + MnO_2$

(*a*) 27.5 (*b*) 55 (*c*) 110 (*d*) 165

29. Equivalent weight of oxidising agent will be-

$2H_2 + O_2 \rightarrow 2H_2O$

(*a*) 1 (*b*) 32 (*c*) 2 (*d*) 8

Multiconcept MCQs

1. For the reaction, $N_2 + 3H_2 \rightarrow 2NH_3$; if E_1 & E_2 are the equivalent masses of NH_3 & N_2; then $E_1 - E_2$ is:

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

2. What mass of HNO_3 is required to make 1 ltr. of 2N solution to be used as oxidising agent in the reaction?

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

(*a*) 63 gm (*b*) 42 gm (*c*) 84 gm (*d*) 126 gm

3. If equal volumes of 1 M $KMnO_4$ & 1 M $K_2Cr_2O_7$ are allowed to oxidise Fe^{2+} to Fe^{3+} in acidic medium, volume of oxidant required for 1 mole of Fe^{2+} will be:

(*a*) $V_{KMnO_4} > V_{K_2Cr_2O_7}$ (*b*) $V_{K_2Cr_2O_7} = V_{KMnO_4}$

(*c*) $V_{K_2Cr_2O_7} > V_{KMnO_4}$ (*d*) None of these

4. In basic medium, CrO_4^{2-} oxidises $S_2O_3^{2-}$ to form SO_4^{2-} and itself changes to $Cr(OH)_4^-$. How many mL of 0.154 M CrO_4^{2-} are required to react with 40 mL of 0.246 M $S_2O_3^{2-}$?

(*a*) 200 mL (*b*) 156.4 mL

(*c*) 170.4 mL (*d*) 190.4 mL

5. Which of the following is/are redox reaction (s)?

(*a*) $SO_2 + 2H_2S \hookrightarrow 2H_2O + 3S$

(*b*) $2BaO + O_2 \rightarrow 2BaO_2$

(*c*) $2KClO_3 \rightarrow 2KCl + 3O_2$

(*d*) All are correct

6. 0.5 g of fuming H_2SO_4 (oleum) is diluted with water. This solution is completely neutralized by 26.7 mL of 0.4 N KOH. The percentage of free SO_3 in the sample is

(*a*) 30.6% (*b*) 40.6% (*c*) 20.6% (*d*) 50%

7. Standard electrode potentials of redox couples A^{2+}/A, B^{2+}/B, C^{2+}/C and D^{2+}/D are 0.3 V, –0.5 V, –0.75 V and 0.9 V respectively. Which of these is best oxidising agent and reducing agent respectively?

(*a*) D^{2+}/D and B^{2+}/B

(*b*) B^{2+}/B and D^{2+}/D

(*c*) D^{2+}/D and C^{2+}/C

(*d*) C^{2+}/C and D^{2+}/D.

8. Volume of 0.5 M HCl required for complete reaction of 10 g equimolar mixture of Na_2CO_3 (molar mass = 106 g) and $NaHCO_3$ (molar mass = 84 g) is

(*a*) 135.8 mL (*b*) 315.8 mL

(*c*) 831.5 mL (*d*) 513.8 mL

9. What volume of 0.1 M $KMnO_4$ is needed to oxidize 100 mg of FeC_2O_4 in acid solution?

(*a*) 4.1 mL (*b*) 8.2 mL (*c*) 10.2 mL (*d*) 4.6 mL

10. A, B and C are three elements forming part of a compound in oxidation states of +2, +5 and –2 respectively. What could be the compound?

(*a*) $A_2(BC)_2$ (*b*) $A_2(BC_4)_3$

(*c*) $A_3(BC_4)_2$ (*d*) ABC

11. Which one is correctly matched?

	Substance	O.N. of S
(*a*)	H_2S	+2
(*b*)	H_2SO_5	+4
(*c*)	H_2SO_4	+4
(*d*)	$Na_2S_4O_6$	+2.5

12. In the following reaction

$O_3 + 6I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$

equivalent weight of O_3 (with molecular weight M) is-

(*a*) $\frac{M}{2}$ (*b*) $\frac{M}{4}$

(*c*) $\frac{M}{24}$ (*d*) $\frac{M}{6}$

13. Identify the disproportionation reaction:

(*a*) $H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

(*b*) $S_8(s) + 12OH^-(aq) \rightarrow 4S^{2-}(aq) + 2S_2O_3^{2-}(aq)$

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

(*d*) All of these

14. 2 mole of N_2H_4 loses 20 moles of electrons to form a compound Y. Assuming that all nitrogen appears in the new compound, if there is no change in O.N. of hydrogen, the O.N. of nitrogen in Y is

(*a*) + 3 (*b*) – 3

(*c*) + 1 (*d*) + 5

15. In the following reaction (unbalanced), equivalent weight of As_2S_3 is related to molecular weight M by :

$As_2S_3 + H^+ + NO_3^- \rightarrow NO + H_2O + AsO_4^{3-} + SO_4^{2-}$

(a) $\frac{M}{2}$ (b) $\frac{M}{4}$

(c) $\frac{M}{28}$ (d) $\frac{M}{24}$

16. Consider the following statement(s)

(A) Iron reacts with dilute H_2SO_4 to evolve $H_2(g)$ but Ag does not.

(B) Copper dissolves in dilute HNO_3 but not in dilute HCl.

(C) Amongst halogens, fluorine is the best oxidant because it is the most electronegative element known.

Choose the correct statement(s)

(a) A & B (b) B & C

(c) A & C (d) A, B & C

17. An element A in a compound ABD has oxidation number A^{n-}. It is oxidised by $Cr_2O_7^{2-}$ in acidic medium. In the experiment 1.68×10^{-3} mole of $K_2Cr_2O_7$ were used for 3.26×10^{-3} mole of ABD. The new oxidation number of A after oxidation is:

(a) 3 (b) 3 – n

(c) n – 3 (d) +n

18. 2 mole of $FeSO_4$ are oxidised by 'X' mole of $KMnO_4$ whereas 2 mole of FeC_2O_4 are oxidised by 'Y' mole of $KMnO_4$. The ratio of 'X' and 'Y' is:

(a) 1 : 3 (b) 1 : 2

(c) 1 : 4 (d) 1 : 5

19. Which of the following reaction is an example of redox reaction?

(a) $CH_2 = CH_2(g) + H_2(g) \rightarrow H_3C - CH_3(g)$

(b) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

(c) $XeF_6(g) + 3H_2O(l) \rightarrow XeO_3 + 6HF$

(d) All of these

20. In the balanced chemical reaction,

$IO_3^- + pI^- + qH^+ \rightarrow rH_2O + sI_2$

p, q, r and s respectively correspond to

(a) 5, 6, 4, 3 (b) 5, 6, 3, 3

(c) 4, 3, 5, 6 (d) 4, 5, 6, 4

NEET Past 10 Years Questions

1. Which of the following reactions is the metal displacement reaction? Choose the right option. *(2021)*

(a) $Cr_2O_3 + 2Al \xrightarrow{\Delta} Al_2O_3 + 2Cr$

(b) $Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$

(c) $2Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2\uparrow$

(d) $2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$

2. What is the change in oxidation number of carbon in the following reaction? *(2020)*

$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$

(a) 0 to + 4 (b) –4 to + 4

(c) 0 to –4 (d) +4 to + 4

3. The oxidation number of the underlined atom in the following species *(2020 Covid Re-NEET)*

Identify the incorrect option

(a) $\underline{Cl}O_3^-$ is +5 (b) $K_2\underline{Cr}_2O_7$ is +6

(c) $H\underline{Au}Cl_4$ is +3 (d) $Cu_2\underline{O}$ is –1

4. Which of the following reactions are disproportionation reaction? *(2019)*

A. $2Cu^+ \longrightarrow Cu^{2+} + Cu^0$

B. $3MnO_4^{2-} + 4H^+ \longrightarrow 2MnO^-_4 + MnO_2 + 2H_2O$

C. $2KMnO_4 \xrightarrow{\Delta} K_2MnO_4 + MnO_2 + O_2$

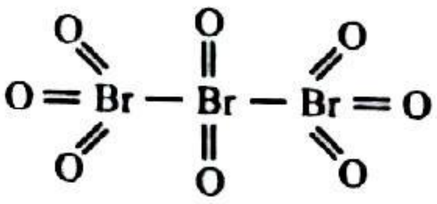
D. $2MnO_4^- + 3Mn^{2+} + 2H_2O \longrightarrow 5MnO_2 + 4H^{\oplus}$

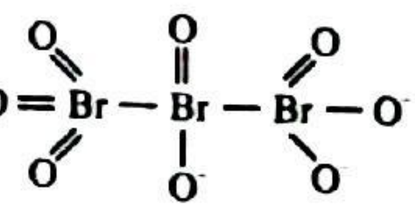
Select the correct option from the following

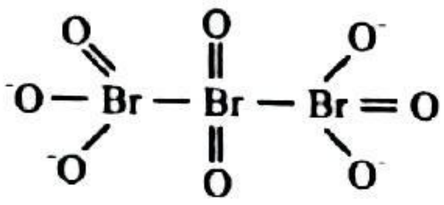
(a) (A) and (B) only (b) (A), (B) and (C)

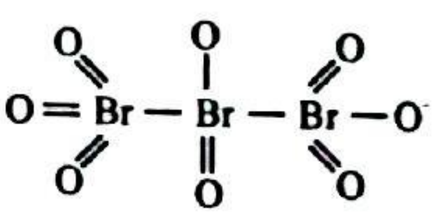
(c) (A), (C) and (D) (d) (A) and (D) only

5. The correct structure of tribromooctaoxide is *(2019)*

(a) 

(b) 

(c) 

(d) 

6. For the redox reaction

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

The correct coefficients of the reactants for the balanced equation are *(2018)*

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

7. Hot concentrated sulphuric acid is a moderately strong oxidising agent. Which of the following reactions does not show oxidising behaviour? *(2016 - I)*

(*a*) $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$

(*b*) $CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$

(*c*) $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

(*d*) $3S + 2H_2SO_4 \rightarrow 3SO_2 + 2H_2O$

8. Assuming complete ionisation, same moles of which of the following compounds will require the least amount of acidified $KMnO_4$ for complete oxidation? *(2015 Re)*

(*a*) $Fe(NO_2)_2$ (*b*) $FeSO_4$

(*c*) $FeSO_3$ (*d*) FeC_2O_4

9. 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: *(2014)*

(*a*) +3 (*b*) +6

(*c*) −10 (*d*) +5

10. A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction which element undergoes maximum change in the oxidation number? *(2012 Pre)*

(*a*) C (*b*) S

(*c*) H (*d*) Cl

11. In which of the following compounds, nitrogen exhibits highest oxidation state? *(2012 Pre)*

(*a*) NH_2OH (*b*) N_2H_4

(*c*) NH_3 (*d*) N_3H

ANSWER KEY

Topicwise Questions

1. (*b*)	2. (*a*)	3. (*a*)	4. (*a*)	5. (*b*)	6. (*d*)	7. (*a*)	8. (*c*)	9. (*a*)	10. (*c*)
11. (*a*)	12. (*d*)	13. (*a*)	14. (*b*)	15. (*a*)	16. (*c*)	17. (*b*)	18. (*a*)	19. (*d*)	20. (*c*)
21. (*a*)	22. (*b*)	23. (*c*)	24. (*c*)	25. (*c*)	26. (*b*)	27. (*a*)	28. (*c*)	29. (*c*)	30. (*c*)
31. (*d*)	32. (*c*)	33. (*a*)	34. (*c*)	35. (*c*)	36. (*b*)	37. (*d*)	38. (*a*)	39. (*a*)	40. (*a*)
41. (*b*)	42. (*a*)	43. (*b*)	44. (*a*)	45. (*c*)	46. (*b*)	47. (*d*)	48. (*b*)	49. (*c*)	50. (*d*)
51. (*d*)	52. (*c*)	53. (*c*)	54. (*d*)	55. (*d*)	56. (*b*)	57. (*a*)	58. (*c*)	59. (*b*)	60. (*c*)
61. (*a*)	62. (*c*)	63. (*c*)	64. (*a*)	65. (*a*)	66. (*d*)	67. (*a*)			

Learning Plus

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

Multiconcept MCQs

1. (*a*)	2. (*b*)	3. (*a*)	4. (*c*)	5. (*d*)	6. (*c*)	7. (*c*)	8. (*b*)	9. (*a*)	10. (*c*)
11. (*d*)	12. (*d*)	13. (*d*)	14. (*a*)	15. (*c*)	16. (*a*)	17. (*b*)	18. (*a*)	19. (*a*)	20. (*b*)

NEET Past 10 Years Questions

1. (*a*)	2. (*b*)	3. (*d*)	4. (*a*)	5. (*a*)	6. (*b*)	7. (*b*)	8. (*b*)	9. (*b*)	10. (*d*)
11. (*d*)									

Solution

Topicwise Questions

1. (*b*) Oxidation number increases and number of electrons decreases in oxidation process.

2. (*a*) F_2, being most electronegative, is the most powerful oxidising agent with standard oxidation potential 2.87 V.

3. (*a*) In the reaction

$$HOCl + \overset{-1}{H_2O_2} \xrightarrow{H^-} H_2O^{\cdot} + Cl^- + O_2$$

↑ in O.N.

Oxidation number of oxygen increases from -1 to 0. Hence, H_2O_2 acts as reducing agent.

4. (*a*)

↓ in O.N. (reduction)

$$\overset{-1}{Ag_2O} + \overset{-1}{H_2O_2} \longrightarrow \overset{0}{2Ag} + H_2O + O_2^0$$

↑ in O.N. (oxidation)

The oxidation number of O in H_2O_2 increases. Hence, H_2O_2 acts as a reducing agent as it is undergoing oxidation.

5. (*b*)

Reduction

$$\overset{2+}{Cu} + \overset{0}{Ag} \longrightarrow \overset{0}{Cu} + \overset{+1}{Ag^{\cdot}}$$

Oxidation

So; oxidation-half reaction is $Ag \rightarrow Ag^{\cdot}$

6. (*d*) $I_2 + 10H\overset{+5}{N}O_3 \rightarrow 2HIO_3 + 10\overset{+4}{N}O_2 + 4H_2O$

In this reaction, HNO_3 is acting as an oxidising agent because oxidising state of N decrease from + 5 to +4

(b) $3Cu + 8H\overset{+5}{N}O_3 \rightarrow 3Cu(NO_3)_2 + 2\overset{+4}{N}O + 4H_2O$

It is a reaction in which HNO_3 is acting as an oxidising agent because oxidation number of nitrogen is decreasing.

(c) $4Zn + 10H\overset{+5}{N}O_3 \rightarrow 4Zn(NO_3)_2 + \overset{-3}{N}H_4NO_3 + 3H_2O$

∴ In this reaction also, HNO_3 act as oxidising agent.

(d) $2H\overset{+5}{N}O_3 + P_2O_5 \rightarrow 2HPO_3 + \overset{+5}{N}_2O_5$

HNO_3 does not act as an oxidising agent because P_2O_5 is a dehydrating agent. So, in this reaction also, there is no change in oxidation number of N.

7. (*a*) In $MnSO_4$, oxidation state of Mn = +2. In Mn_2O_3, oxidation state of Mn = +3. There is an increase of oxidation state by 1 when $MnSO_4$ changes to Mn_2O_3. Hence; its equivalent mass is equal to its molar mass.

$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{\text{Oxidation state}}$$

8. (*c*) Oxidation is the addition of electronegative element or removal of electropositive element or increase in oxidation number.

$Cl_2 + 2KI \rightarrow 2KCl + I_2$

In this reaction, the underlined species is oxidised as it involve the removal of electropositive element, i.e., $K^{\cdot}$

9. (*a*) The reaction which involves oxidation, requires an oxidising agent. $S_2O_3^{2-} \rightarrow S_4O_6^{2-}$

S is undergoing oxidation. Oxidation state of S changes from 2 to 2.5 and hence this reaction requires an oxidising agent.

10. (*c*) The species in which an element has oxidation number lying in between its minimum and maximum values can act as oxidant and reductant both.

Oxidation state of O in species H_2O_2	Minimum oxidation state of O	Maximum oxidation state of O
–1 in H_2O_2	–2	0

Thus: hydrogen peroxide (H_2O_2) can act oxidising as well as a reducing agent.

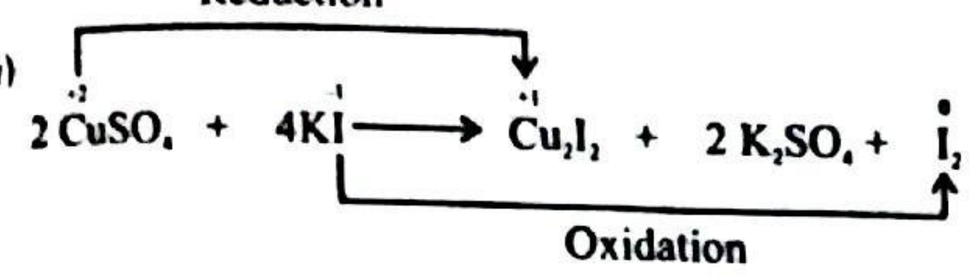
H_2O_2 as a reducing agent.

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

H_2O_2 as an oxidising agent.

$PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$

11. (*a*)



12. (*d*) Theory based.

13. (*a*) Meq. of oxalic acid = Meq. of $KMnO_4$

$$V \times 0.1 = \frac{250 \times 8}{100 \times 31.6} \times 1000 = 6.3 \text{ litre}$$

14. (*b*) $6e^- + (N^0)_2 \rightarrow 2N^{-3}$

$$\therefore E_{N_2} = \frac{28}{6}; E_{NH_3} = \frac{17}{3}$$

15. (*a*) Meq. of lime stone = Meq. of CaC_2O_4 = Meq. of $KMnO_4$ = Meq. of CaO

$$\therefore \quad 40 \times 0.250 = \frac{w}{56/2} \times 1000$$

$$\therefore \quad w_{CaO} = 0.28$$

$$\therefore \text{ \% of CaO} = \frac{0.28 \times 100}{0.518} = 54\%$$

16. (*c*) 5 mole I^- gives 3 mole I_2.

17. (*b*) $3Fe^0 \rightarrow Fe_3^{+(8/3)} + 8e^-$

$$\therefore E = \frac{M}{8/3} = \frac{56 \times 3}{8} = 21$$

18. (*a*) The oxidation number of an element is zero. Therefore; its covalency is 2 as oxidation number of S in S_8 is zero. Further in the S_8 molecule, each sulphur atom is connected to two other sulphur atoms, therefore covalency is 2.

19. (*d*) Let the oxidation state of N in HNO_3 is x.

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

$x + 1 - 6 = 0$

$x = +5$

20. (*c*) NO_3^- has a uni-negative charge.

$\therefore$ In $Fe(NO_3)_2$, Fe has oxidation number + 2

$C_2O_4^-$ has bi-negative charge.

$\therefore$ In FeC_2O_4, Fe has oxidation number + 2

In $FeCl_3$, Fe has + 3 O.S

$\therefore$ In Mohr's salt also, Fe has +2 oxidation number.

21. (*a*) Let the oxidation state of P in the given compounds be x.

$H_4P_2O_5 : 2x + 4 - 10 = 0$, $H_4P_2O_6 : 2x + 4 - 12 = 0$

$2x = 6$ $\qquad$ $2x = 8$

$x = +3$ $\qquad$ $x = +4$

$H_4P_2O_7 : 2x + 4 - 14 = 0$

$2x = 10$

$x = +5$

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

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

$2 + x - 2 = 0$

$x = 0$

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

$2 + y - 4 = 0$

$y = +2$

23. (*c*) Let the oxidation number of Co be x.

In $K[Co(CO)_4]$, $1 + x + 0 = 0$ [as CO is neutral]

$x = -1$

24. (*c*) Oxidation state of I in IO_3^-

$x + (-2)3 = -1 \Rightarrow x = +5$

Oxidation state of I in IO_4^-

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

Oxidation state of I in KI

$+1 + x = 0 \Rightarrow x = -1$

Oxidation state of I in I_2 $x = 0$

Hence; the oxidation number of iodine in IO_3^-, IO_4^-, KI and I_2 are respectively + 5, + 7, –1 and 0

25. (*c*) Acidified $K_2Cr_2O_7$ solution oxidises SO_2 into $Cr_2(SO_4)_3$

$$3\overset{+4}{S}O_2 + K_2Cr_2O_7 + H_2SO_4 \rightarrow K_2SO_4 + Cr_2\left(\overset{+6}{S}O_4\right)_3 + H_2O$$

Hence; the oxidation state of sulphur change from +4 to +6

26. (*b*) The highest oxidation state is achieved by $(n-1)\ d^5\ ns^2$ electronic configuration. This shows the maximum oxidation state, i.e., +7

27. (*a*) In $Ba(H_2PO_2)_2$, let the oxidation state of P be x

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

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

$2 + 4 + 2x - 8 = 0$

$2 + 2x - 4 = 0$

$2x = 2 \Rightarrow x = +1$

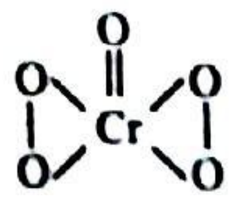
28. (*c*) $Mn^{2+} + NO_3^- + OH^- \rightarrow \overset{(+6)}{Mn}O_4^{2-} + H_2O + NO$

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

29. (*c*) When H_2O_2 is added to an acidified solution of a $Cr_2O_7^{2-}$, a deep blue coloured complex, chromic peroxide, CrO_5 [or $CrO(O_2)_2$] is formed.

$Cr_2O_7^{2-} + 2H^+ + 4H_2O_2 \rightarrow 2CrO(O_2)_2 + 5H_2O$

This deep blue coloured complex has the following structure:



Oxidation state of Cr is +6 due to the presence of two peroxide linkages, which can be calculated as:

In $CrO(O_2)_2$, let the oxidation state of Cr be x

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

$x - 6 = 0$

$x = +6$

30. (*c*) The following reaction occurs when a mixture of potassium chlorate, oxalic acid and sulphuric acid is heated:

$$\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} + \overset{+1}{H_2}\overset{+3}{C_2}\overset{-2}{O_4} + \overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4} \rightarrow \overset{+1}{K_2}\overset{+6}{S}\overset{-2}{O_4} + \overset{+1}{K}\overset{-1}{Cl} + \overset{+4}{C}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

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

31. (*d*) $\overset{2.5}{Na_2S_4O_6} : \overset{+6}{H_2S_2O_7} : \overset{+2}{Na_2S_2O_3} : \overset{0}{S_8}$

$\overset{+6}{H_2SO_4} : \overset{+4}{SO_2} : \overset{-2}{H_2S} : \overset{+7}{H_2S_2O_8}$

$\overset{+4}{SO_2} ; \overset{+6}{SO_4^{2-}} ; \overset{+4}{SO_3^{2-}} : \overset{+6}{HSO_4^-}$

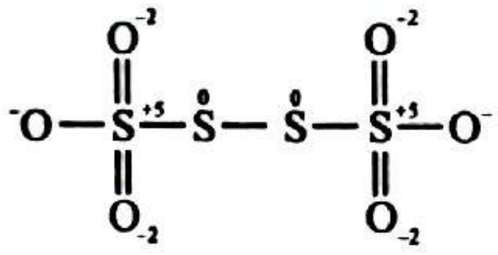
$\overset{+6}{H_2SO_4}$; $\overset{+4}{H_2SO_3}$; $\overset{+2}{SCl_2}$; $\overset{-2}{H_2S}$

32. (*c*) The tetrathionate anion, $S_4O_6^{2-}$ is a sulphur oxoanion derived from the compound tetrathionic acid, $H_2S_4O_6$ anion $S_4O_6^{2-}$ can be viewed as the product formed from reaction between lewis base, S_2^{2-} and lewis acid, SO_3

$$S_2^{2-} + 2SO_3 \rightarrow S_4O_6^{2-}$$

Two of the sulphur atoms are present in the oxidation state of zero and two are in oxidation state +5.

The structure of $S_4O_6^{2-}$



Since; each of the terminal sulphur atom is connected to two oxygen atom by a double bond and to one oxygen atom by a single bond, the oxidation state of these terminal sulphur atoms is +5. Since; two central sulphur atoms are linked to each other by a single bond, and the sulphur is further attached to similar species on either side, the electron pair forming S-S bond remains in the centre and hence; each of the two central sulphur atoms has on oxidation state of zero. Thus; the average oxidation state of the four sulphur atom is

$(2 \times 5 + 2 \times 0)/4 = 2.5$

33. (*a*) Bromine has zero oxidation state because it is in the free state.

$Br_2^0 \rightarrow BrO_3^-$

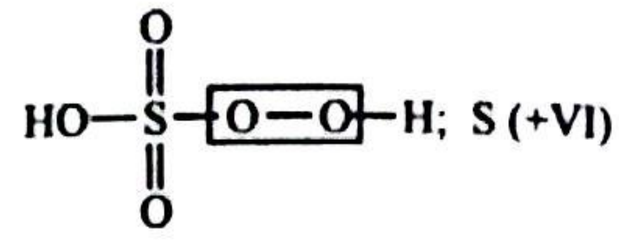
$x + (-2 \times 3) = -1$

$x + (-6) = -1$

$x = +6 = +5$

So: oxidation number changes from 0 to + 5

34. (*c*) H_2SO_5 (Peroxomonosulphuric acid)

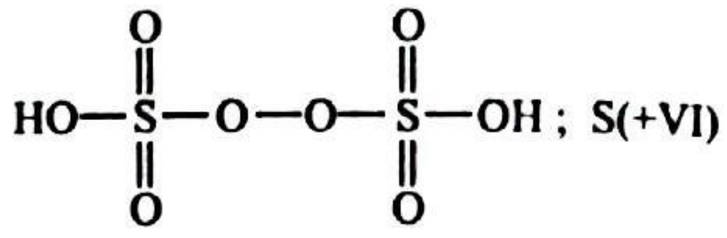


$= 2 \times (+1) + x + 3(-2) + 2(-1) = 0$

$2 + x - 6 - 2 = 0$

$x = +6$

$H_2S_2O_8$ (peroxodisulphuric acid)



The oxidation number of these two oxygen atoms will be –1 each because it contains peroxide linkage.

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

$2 + 2x - 12 - 2 = 0$

$2x = 12$

$x = +6$

35. (*c*) In Fe_3O_4, let the oxidation number of Fe = x

$3x + 4(-2) = 0$

$$x = \frac{8}{3}$$

36. (*b*) The reaction of aqueous $KMnO_4$ with H_2O_2 in acidic medium is

$3H_2SO_4 + 2KMnO_4 + 5H_2O_2 \rightarrow 5O_2 + 2MnSO_4 + 8H_2O + K_2SO_4$

In the above reaction, $KMnO_4$ oxidises H_2O_2 to O_2 and itself (MnO_4^-) gets reduced to Mn^{2+} ion in acidic conditions.

37. (*d*) a. Oxidation state of C in CO_2

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

b. Oxidation state of C in CF_4

$x + 4(-1) = 0 \Rightarrow x = +4$

c. Oxidation state of C in CCl_4

$x + 4(-1) = 0 \Rightarrow x = +4$

d. Oxidation state of C in CH_4

$x + 4(+1) = 0 \Rightarrow x = -4$

38. (*a*) $H_4P_2O_7$, let the oxidation number of P be x

$2x + 4 - 14 = 0$

$\Rightarrow 2x - 10 = 0$

$\Rightarrow x = +5$

39. (*a*) The oxidation number of nitrogen in different compounds are

a. $Mg_3N_2 - 3$ b. $NH_2OH - 1$

c. $(N_2H_5)_2SO_4 - 2$ d. $[Co(NH_3)_5Cl]Cl_2$ $^{+3}$

40. (*a*) N_3H (hydrazoic acid) let the oxidation state of N be x

$3(x) + 1 = 0$

$3x = -1$

$x = -1/3$

41. (*b*) Hydrogen show electronegative nature with metals.

42. (*a*) O.S of P is +3 in H_3PO_3

43. (*b*) Br_2 changes from 0 to –1 and +5

44. (*a*) Oxidation state remain same of the element on both side.

45. (*c*) It is displacement reaction

46. (*b*) F only show –1 oxidation state

47. (*d*) Here 'CO' is also neutral.

48. (*b*) Because H_2O accept electrons.

49. (*c*) $1 + (-2) + x = 0$

50. (*d*) All the given compounds can gain and loss electrons.

51. (*d*) Learn from theory.

52. (*c*) S-has 6-vallence electrons its show maximum +6 oxidation state.

53. (*c*) $x + 3 \times (-1) = 0 \Rightarrow x = +3$

54. (*d*) The balanced reaction is as follows :

$$8\,KMnO_4 + 3NH_3 \rightarrow 8MnO_2 + 3KNO_3 + 5KOH + 2H_2O$$

55. (*d*) Given, $NO_3^- \rightarrow N_2H_4$

The balanced reaction is as

$$2NO_3^- + 16H^+ + 14e^- \rightarrow N_2H_4 + 6H_2O$$

$\therefore$ Number of electrons involved in the reduction of two nitrate ions = $14e^-$

$\therefore$ One nitrate ion will involve = $\frac{14}{2}e^- = 7e^-$

56. (*b*) Balanced equation is

$$BiO_3^- + 6H^+ + 2e^- \rightarrow Bi^{+3} + 3H_2O$$

$\therefore$ The value of x is 2

57. (*a*) The balance chemical reaction is as

$$3UO^{2+} + CrO_7^{2-} + 8H^+ \rightarrow 3UO_2^{2+} + 2Cr^{3+} + 4H_2O$$

58. (*c*) The balanced chemical reaction is

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

Thus; when KI reacts with 1 mole of potassium dichromate, 3 moles of I_2 are evolved.

59. (*b*) In alkaline medium, following reaction occurs,

$$2KMnO_4 + H_2O \xrightarrow{\text{alkaline}} 2MnO_2 + 2KOH + 3[O]$$

$$2\overset{+7}{Mn}O_4^- + H_2O \xrightarrow{\text{alkaline}} 2\overset{+4}{Mn}O_2 + 2OH^- + 3[O]$$

$$\text{So; equivalent mass} = \frac{\text{Molar mass}}{3}$$

60. (*c*) Balanced equation is

$$2IO_3^- + 5SO_2 + 4H_2O \rightarrow I_2 + 5SO_4^{2-} + 8H^+$$

The coefficient of SO_2 is 5.

61. (*a*) $I^- + \overset{+5}{I}O_3^- + H^+ \rightarrow \overset{0}{I_2} + H_2O$

$$2I^- \rightarrow I_2 + 2e^- \qquad \text{(i)}$$

$$10e^- + 2(\overset{+5}{I}O_3^-) \rightarrow \overset{0}{I_2} \qquad \text{(ii)}$$

Multiplying eq (i) by 5 and then adding to eq (ii), in order to cancel the electrons, we get

$$10I^- + 2IO_3^- \rightarrow 6I_2$$

To balance O-atom, added $6H_2O$ molecules on RHS and $12H^+$ on L.H.S, then

$$10I^- + 2IO_3^- + 12H^+ \rightarrow 6I_2 + 6H_2O$$

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

62. (*c*) $ClO_2 \rightarrow Cl^-$

$$ClO_2 + 2H_2O + 5e^- \rightarrow Cl^- + 4OH^-$$

$$H_2O_2 \rightarrow O_2$$

$$H_2O_2 + 2OH^- \rightarrow O_2 + 2H_2O + 2e^-$$

$$[ClO_2 + 2H_2O + 5e^- \rightarrow Cl^- + 4OH^-] \times 2$$

$$[H_2O_2 + 2OH^- \rightarrow O_2 + 2H_2O + 2e^-] \times 5$$

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

$2ClO_2 = 5H_2O \Rightarrow ClO_2 = 2.5\,H_2O_2$

Hence one mole of ClO_2 oxidises 2.5 moles of H_2O_2

63. (*c*) $4Zn + NO_3^- + 7H_2O \rightarrow 4Zn^{2+} + NH_4^+ + 10\,OH^-$

64. (*a*) In which same.

65. (*a*) Balance the reaction oxidation state of element convert into different.

66. (*d*) Magnesium is more electropositive than zinc.

67. (*a*) Since the O.N. of 'Fe' changes by 3, eq. wt. of FeC_2O_4 in the change is M/3.

Learning Plus

1. (*d*) Following are the examples of redox reaction

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

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

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

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

2. (*d*) $Fe^{3+}/Fe^{2+} = +0.77$

$I_2(s)/I^- = +0.54$;

$Cu^{2+}/Cu = +0.34$; $Ag^+/Ag = 0.80$ V

Ag^+ is the strongest oxidising agent as E° of the redox couple Ag^+/Ag is the most positive, i.e., 0.80 V therefore, Ag^+ is the strongest oxidising agent.

3. (*d*) Given that E° values of

$E^\ominus$ values : $Br_2/Br^- = +1.90$

$Ag^+/Ag(s) = +0.80$

$Cu^{2+}/Cu(s) = +0.34$

$I^-/I_2(s) = +0.54$

$Br^-/Br_2 = -1.90$ V

The E° values show that copper will reduce Br_2, if the E° of the following redox reaction is positive

$Cu + Br_2 \rightarrow CuBr_2$

Now, $Cu \rightarrow Cu^{2+} + 2e$; $E^\circ = -0.34$ V

$$\frac{Br_2 + 2e^- \rightarrow 2Br^- ; E = +1.09V}{Cu + Br_2 \rightarrow CuBr_2 ; \partial E = +0.75V}$$

Since, E° of this reaction is positive, therefore, Cu can reduce Br_2. While other reaction will give negative value.

4. (*d*) $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$; $E^\circ = +0.77$ V

$2I^- \rightarrow I_2 + 2e^-$; $E^\circ = -0.54$ V (sign of E° is reversed)

$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$; $E^\circ_{cell} = +0.23$ V

This reaction is feasible since E°_{cell} is positive

(b) $Cu \rightarrow Cu^{2+} + 2e^-$; $E = -0.34$ V

(sign of E° has been reversed)

$2Ag^+ + 2e^- \rightarrow 2Ag$; $E^\circ = 0.80$ V

$Cu + 2Ag^+ \rightarrow 2Cu^{2+} + 2Ag$; $E^\circ = +0.46$ V

This reaction is feasible since E°_{cell} is positive.

(*c*)
$$\frac{2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+} ; E^\circ = +0.77\,V}{Cu \rightarrow Cu^{2+} + 2e^- ; E^\circ = -0.34\,V}$$
$$2Fe^{3+} + Cu \rightarrow 2Fe^{2+} + Cu^{2+} ; E^\circ = +0.43V$$

(sign of E° is reversed)

This reaction is feasible since E°_{cell} is positive.

(*d*) $Ag \rightarrow Ag^+ + e^-$; $E^\circ = 0.80$ V (sign of E° is reversed)

$Fe^{3+} + e^- \rightarrow Fe^{2+}$; $E^\circ = +0.77$V

$Ag + Fe^{3+} \rightarrow Ag^+ + Fe^{2+}$; $E^\circ = -0.03$V

This reaction is not feasible since E°cell is negative.

5. (*a*) $2\overset{+2}{S_2}\overset{+2}{O_3^{2-}}(aq) + \overset{0}{I_2}(s) \rightarrow \overset{2.5}{S_4}O_6^{2-}(aq) + 2I^-(aq)$

$\overset{+2}{S_2}\overset{+2}{O_3^{2-}}(aq) + 2\overset{0}{Br_2}(l) + 5H_2O(l) \rightarrow$

$2SO_4^{2-}(aq) + 4Br^-(aq) + 10H^+(aq)$

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.

6. (*a*) Oxidation number of hydrogen is always +1 is a wrong rule since, it is +1 in hydrogen halides, – 1 in hydrides and zero in H_2 molecule.

All the other three statements (*b*), (*c*) and (*d*) are correct.

7. (*b*) NH_4NO_3 is actually NH_4^+ and NO_3^-. It is an ionic compound Nitrogen in the two species have different as oxidation no. shown below

Let, oxidation number of N in $\overset{-3}{N}H_4$ is x.

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

or $x + 4 = +1$ or $x = -3$

Let, oxidation number of N in $\overset{+5}{N}O_3^-$ is x

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

or $x - 6 = -1$ or $x = +5$

8. (*a*) 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 O. N. of central atom increases from left to right. Therefore, option (a) is correct.

9. (*d*) Highest oxidation number of any transition element = (n – 1) d electrons + ns electrons. Therefore, large 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

10. (*d*) Disproportionation reactions are reactions in which the same substance is oxidised as well as reduced are called Writing the O.N. of each element above its symbol in the given reactions.

(a) $\overset{-4+1}{C}H_4 + 2\overset{0}{O_2} \rightarrow \overset{+4-2}{C}O_2 + 2\overset{+2}{H_2}\overset{-2}{O}$

(b) $\overset{-4+1}{C}H_4 + 4\overset{0}{Cl_2} \rightarrow \overset{+4-1}{C}Cl_4 + 4\overset{+1-1}{H}Cl$

(c) $2\overset{0}{F_2} + 2\overset{-2+1}{O}H \rightarrow 2\overset{-1}{F^-} + \overset{+2}{O}\overset{-1}{F_2} + \overset{+1}{H_2}\overset{-2}{O}$

(d) $2\overset{+4-2}{N}O_2 + 2OH^- \rightarrow \overset{+3-2}{N}O_2^- + \overset{+5-2}{N}O_3^- + \overset{+1}{H_2}\overset{-2}{O}$

Thus, in reaction (*d*), N is both oxidised as well as reduced since the O.N. of N increases from + 4 in NO_2 to + 5 in NO_3^- and decreases from + 4 in NO_2 to + 3 in NO_2^-

11. (*c*) F always shows an oxidation number of –1 as being the most electronegative element, F can only be reduced Further, due to the absence of d-orbitals it cannot be oxidised and hence it does not show positive oxidation numbers.

In other words, F cannot be oxidised as well as reduced simultaneously and hence does not show disproportionation reactions.

12. (*a*) $30H^+ + 5MnO_4^- + 5Fe(HC_2O_4)_2 \rightarrow 5Mn^{2+} + 5Fe^{3+} + 20CO_2 + 20H_2O$

From here it is evident that 1 mole of $5MnO_4^-$ is required to react with 1 mole of $Fe(HC_2O_4)_2$.

$$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$$

$$Fe(HC_2O_4)_2 \rightleftharpoons Fe^{2+} + 2HC_2O_4^-$$

$$2HC_2O_4^- \rightleftharpoons 2H^+ + C_2O_4^{2-}$$

One mole of $Fe(HC_2O_4)_2$ will furnish 2 mole of H^+ which require 1 mole of $Ca(OH)_2$.

It is not a redox reaction it is rather a neutralisation reaction.

13. (*c*) Eq. wt. in disproportionation reaction

$$= \frac{M}{\text{n factor}_1} + \frac{M}{\text{n factor}_2}$$

$$96 = \frac{160}{2} + \frac{160}{\text{n factor}_2} \Rightarrow 16 = \frac{160}{\text{n factor}_2}$$

$\text{n factor}_2 = 10$

which is possible only if the oxidised product is BrO_3^-.

14. (*a*) $KHC_2O_4 \rightleftharpoons K^+ + HC_2O_4^- \rightleftharpoons H^+ + C_2O_4^{2-}$

$KOH \rightleftharpoons K^+ + OH^-$

There is one replaceable H in KHC_2O_4, so its n - factor is 1.

Its eq. wt $= \frac{128}{1} = 128.$

15. (*a*) $H_2O + 3HC_2O_4^- + 2MnO_4^- \rightarrow 2MnO_2 + CO_2 + 5OH^-$

Eq. wt. of $KHC_2O_4 = \frac{\text{M.wt}}{2} = 64$

16. (*b*) $48H^+ + 6MnO_4^- + 5Fe_2(C_2O_4)_3 \rightarrow 6Mn^{2+} + 3CO_2 + 10Fe^{3+} + 24H_2O$

5 moles of ferric oxalate require $MnO_4^- = 6$ moles

1 moles of ferric oxalate require $MnO_4^- = \frac{6}{5} = 1.2$

17. (*d*) There is peroxy linkage in $S_2O_8^{2-}$

18. (*b*) $2Cr(s) + 2OH^-(aq) + 6H_2O \rightarrow 2Cr(OH)_4^-(aq) + 3H_2(g)$

$2 + 2 + 6 + 2 + 3 = 15$

n-factor = 5

19. (*c*) $As_2O_3 + MnO_4^- \rightarrow 2AsO_4^{3-} + Mn^{2+}$

n-factor = 4

Let, molarity of $KMnO_4$ solution be M

$\therefore$ Eq. of As_2O_3 = Eq. of $KMnO_4$ solution

$$\frac{0.1097}{198} \times 4 = \frac{36.10 \times M \times 5}{1000}$$

$$\left(\text{Equivalent weight } As_2O_3 = \frac{198}{4}\right)$$

Molarity = 0.0122 M

20. (*a*) Meq. of Acid = Meq. of NaOH

$$\frac{0.52}{E} \times 1000 = 100 \times 0.2$$

$\therefore$ E = 26

21. (*c*) Meq. of H_2S = Meq. of Cu^{2+}

$$\therefore \frac{W_{H_2S}}{34/2} \times 1000 = \frac{63.5}{63.5//2} \times 1000$$

$\therefore W_{H_2S} = 34$ g

22. (*a*) Moles of (H^+) = moles of (OH^-) = 1, H_3PO_2.

23. (*b*) Meq. of $NaBrO_3 = 85.5 \times 0.672 = 57.456$

Let weight of $NaBrO_3$ = W

$$\therefore \frac{W}{\text{E.wt}_{NaBrO_3}} \times 6 \times 1000 = 57.456 \text{ (equivalent weight = M/6)}$$

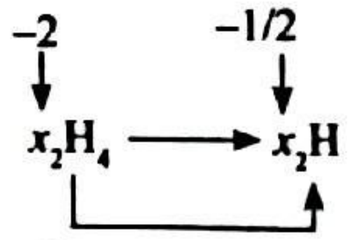
of n-factor = 6

$$\therefore \frac{W}{151} \times 6 \times 1000 = 57.456$$

$$W = \frac{57.456 \times 151}{6 \times 1000} = 1.446 \text{ g}$$

24. (*c*) Green color appears because of $CrCl_3$ in solution

25. (*c*)



Change in O.S. = 1.5

Total electrons based by 1mol = 1.5 × 2 = 3

26. (*b*) O_2PtF_6 [0 = 0 Pt F_6]

or

+1 −1

O_2 — Pt F_6

So [O] have +1/2 O.S.

27. (*c*) Learn P.T.

28. (*a*) $\text{E.wt} = \frac{M}{\text{change in O.S}}$

29. (*d*) Equivalent wt. avg. $O_2 = (32/4) = 8$

Multiconcept MCQs

1. (*a*) Eq. mass of $NH_3 = \frac{17}{3}$

Eq. mass of $N_2 = \frac{28}{6}$

$$E_1 - E_2 = \frac{17}{3} - \frac{28}{6} = \frac{34-28}{6} = \frac{6}{6} = 1$$

2. (*b*) All 8 moles of HNO_3 are not acting as oxidising agent. Just 2 mole are acting as oxidising agent & they themselves have been reduced to NO.

$\Rightarrow$ Gram equivalents of HNO_3 reqd. = 2 × 1 = 2

No. of gm equivalent = no. of moles × n factor

2 = no. of moles × 3

$\frac{2}{3} \times 63 = 42$ g

1 mole of HNO_3 weight = 63g

$\frac{2}{3} \times 63 = 42$ g of HNO_3 weight = $\frac{2}{3} \times 63 = 42$ g

3. (*a*) (1) $8H^+ + MnO_4^- + 5Fe^{2+} \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

(2) $14H^+ + Cr_2O_7^{2-} + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2$

From 2 balanced equations, it is evident that 1 mole of $Cr_2O_7^{2-}$ oxidises 6 moles of Fe^{2+} (and 1 mole of MnO_4^- oxidises 5 moles of Fe^{2+}). So, volume of $KMnO_4$ required is more.

4. (*c*) $40 \times 0.246 \times 8 = V \times 0.154 \times 3$ (Meq. of $S_2O_3^{2-}$ = Meq. of CrO_4^{2-})

$\therefore$ V = 170.4 mL

5. (*d*) O.S. change in the reactions.

6. (*c*) Meq. of H_2SO_4 + Meq.of SO_3 = Meq. of KOH

$$\therefore \frac{(0.5-x)}{98/2}100 + \frac{x}{80/2}100 = 26.7 \times 0.4$$

$\therefore x = 0.103$

$$\therefore \%\text{ of } SO_3 = \frac{0.103}{0.5} \times 100 = 20.6\%$$

7. (*c*) The redox couple with maximum reduction potential will be best oxidising agent and with minimum reduction potential will be best reducing agent.

8. (*b*) $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

Total wt. of Na_2CO_3 and $NaHCO_3$ (as per equations)

= 106 + 84 = 190 g

Moles of HCl used = 3

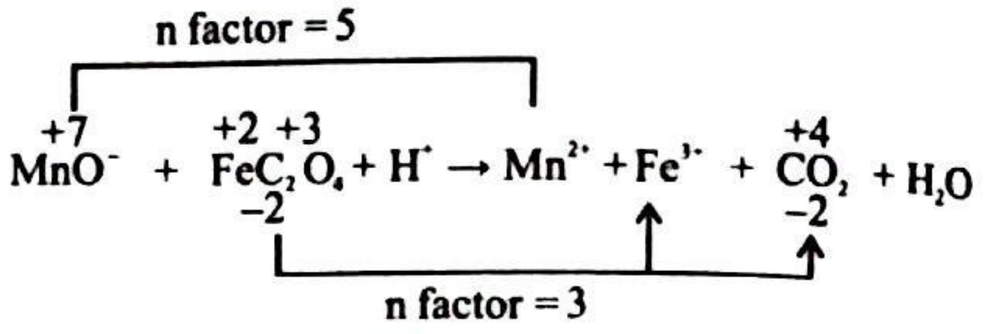
190 g mixture, HCl = 3 mol

10 g mixture, $HCl = \frac{3 \times 10}{190}\text{ mol} = \frac{3}{19}\text{ mol}$

$$(MV)_{HCl} = \frac{3}{19} \Rightarrow V_{HCl(L)} = \frac{3}{19} \times \frac{1}{0.5}$$

$$V_{HCl(mL)} = \frac{3}{19} \times \frac{1}{0.5} \times 1000 = 315.5\,mL$$

9. (*a*) Meq. of $KMnO_4$ = Meq. of FeC_2O_4



$\therefore$ V = 4.1 mL

10. (*c*) $A_3(BC_4)_2$ +6 on A_3 and [+5 −8] × 2 = −6 on $(BC_4)_2$

11. (*d*) Calculate Avg. O.S. of S.

12. (*d*) $E.wt = \frac{M}{\text{change in O.S}}$

13. (*d*) Same O.S. of element convert into different O.S.

14. (*a*) O.N of one 'N' in N_2H_4 = −2

If 2 mole of N_2H_4 loses 20 mol e^-

$\therefore$ 1 mole will lose 10 mol e^-

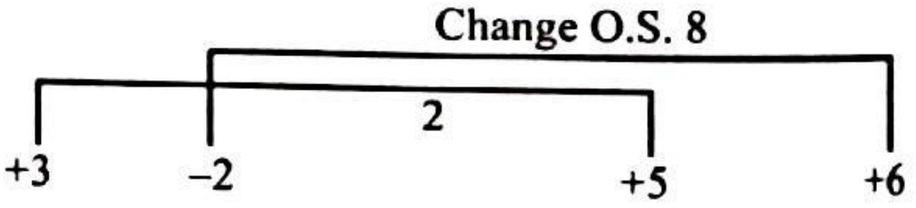
$$\underset{O.S = -2}{N_2H_4} \xrightarrow[-10e^-]{} \text{New compound}$$

$10e^-$ is released from 2 N atom

$1N^-$ atom will release $5e^-$

New O.N of Y = +5 − 2 = +3

15. (*c*)



$n_f = (2 \times 2) + (8 \times 3) = 28$

$E = (M/n_f) = (M/28)$

16. (*a*) Use electrochemical series order.

17. (*b*) $A^{n-} \rightarrow A^{a+} + (a+n)e$

$Cr_2^{6+} + 6e \rightarrow 2Cr^{3+}$

Also, Meq. of A = Meq. of $K_2Cr_2O_7$

$3.26 \times 10^{-3}(a+n) = 1.68 \times 10^{-3} \times 6$

or a + n = 3

$\therefore a = 3 - n$

18. (*a*) $Fe^{2+} \rightarrow Fe^{3+} + e$

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

5 mole $FeSO_4$ = 1 mole $KMnO_4$

'X' = $\frac{2}{5}$ mole

or $Fe^{2+} \rightarrow Fe^{3+} + e$

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

$FeC_2O_4 \rightarrow Fe^{3+} + 2C^{4+} + 3e$

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

5 mole FeC_2O_4 = 3 mole $KMnO_4$

$\therefore$ 'Y' = $\frac{3 \times 2}{5}$

19. (*a*) O.S. change inthe reaction.

20. (*b*) Balance redox reaction.

NEET Past 10 Years Questions

1. (*a*) Aluminium is more electropositive than Cr, so it displaced chromium from Cr_2O_3.

$$Cr_2O_3 + Al \xrightarrow{\Delta} Al_2O_3 + Cr$$

2. (*b*) $CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$

In the given reaction

Let Carbon oxidation number assumes to be x

H oxidation state is +1

In CH_4

$x + 4 \times 1 = 0$

$x = -4$

In CCl_4

Cl oxidation state is –1

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

$x = +4$

Thus, Change in oxidation state of carbon is from –4 to +4

3. (*d*)

(*a*) $\underline{Cl}O_3^-$

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

$x = +5$

(*b*) $K_2\underline{Cr}_2O_7$

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

$x = +6$

(*c*) $H\underline{Au}Cl_4$

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

$x = +3$

(*d*) $\underline{Cu}_2O$

$2(+1) + x = 0$

$x = -2$

4. (*a*)

A. $2\overset{+1}{Cu^+} \longrightarrow \overset{+2}{Cu^{2(+)}} + \overset{0}{Cu}$ } Disproportionation

(red: $Cu^+ \rightarrow Cu$; oxd: $Cu^+ \rightarrow Cu^{2+}$)

B. $3\overset{+6}{Mn}O_4^{2(-)} + 4H^{(+)} \longrightarrow 2\overset{+7}{Mn}O_4^- + \overset{+4}{Mn}O_2 + 2H_2O$ } Disproportionation

(red: $Mn^{6+} \rightarrow Mn^{4+}$; oxd: $Mn^{6+} \rightarrow Mn^{7+}$)

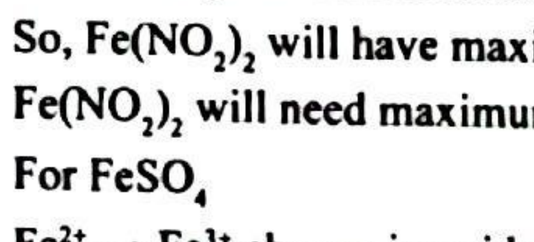
C. $2\overset{+7}{K}\overset{+7}{Mn}O_4^{-2} \longrightarrow K_2\overset{+6}{Mn}O_4 + \overset{+4}{Mn}O_2 + \overset{0}{O_2}$ } $\therefore$

not a disproportionation

D. $2\overset{+7}{Mn}O_4^- + 3\overset{+2}{Mn^{2(+)}} + 2H_2O \longrightarrow 5\overset{+4}{Mn}O_2 + 4H^{\oplus}$ }

5. (*a*) The correct structure of Br_3O_8.

The oxidation no of bromine is $\frac{16}{3}$ in this structure



Tribromooctaoxide

6. (*b*) $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

7. (*b*) $CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$

This reaction is not a oxidation reaction as none of the atom in the reaction is showing any change in the oxidation number. H_2SO_4, here in the reaction is not acting as a reducing nor an oxidising agent.

8. (*b*) $Fe(NO_2)_2 \rightarrow Fe^{3+} + 2NO_3^-$

$Fe^{2+} \rightarrow Fe^{3+}$ change in oxidation state = 1

$2NO_2^- \rightarrow 2NO_3^-$ change in oxidation state = 4

Total change in oxidation state = 5

So, $Fe(NO_2)_2$ will have maximum number of moles.

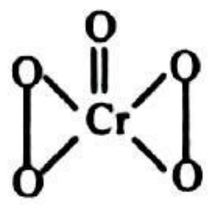
$Fe(NO_2)_2$ will need maximum amount of acidic $KMnO_4$

For $FeSO_4$

$Fe^{2+} \rightarrow Fe^{3+}$ change in oxidation state = 1

$FeSO_4$ will need least amount of $KMnO_4$.

9. (*b*) CrO_5 has a very famous butterfly structure



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

$x = +6$

10. (*d*) ClO_3^- change to Cl^-, oxidation state changes from +5 to –1.

11. (*d*) Let oxidation state of N be *x*

So, oxidation state of N in N_3H = –1/3

$NH_2OH = x + 2 - 1 = -1$

$N_2H_4 = 2x + 4 = -2$

$NH_3 = -3$

$\therefore$ In N_3H, N is present in its highest oxidation state.