8

REDOX REACTION

INTRODUCTION

Many Chemical reactions involved transfer of electrons from one chemical substance to another. These electron-transfer reactions are termed oxidation-reduction or Redox reactions.

OXIDATION AND REDUCTION

There are four concepts for oxidation and reduction reactions

(1) Classical concept (2) Modern concept

(3) Valency concept (4) Oxidation Number concept

1. **Classical concept :** According to this concept, oxidation and reduction can be explained as :

(a) **Oxidation :** Oxidation is a process which involves.

(i) **Addition of Oxygen**

Example :

$2Mg + O_2^* \longrightarrow 2MgO$ (Oxidation of Magnesium)

$C + O_2^* \longrightarrow CO_2$ (Oxidation of Carbon)

$Na_2SO_3 + H_2O_2 \longrightarrow Na_2SO_4 + H_2O$

(Oxidation of Sodium sulphite)

$CH_3CHO + O^* \longrightarrow CH_3COOH$

$2HNO_2 + O_2 \longrightarrow 2HNO_3$

(ii) **Removal of Hydrogen**

Example :

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

(Oxidation of Hydrogen sulphide)

$4HI + O_2^* \longrightarrow 2H_2O + 2I_2$

(Oxidation of Hydrogen iodide)

$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$

(iii) **Addition of Electronegative element**

Example :

$Fe + S^* \longrightarrow FeS$ (Oxidation of Iron)

$SnCl_2 + Cl_2^* \longrightarrow SnCl_4$

(Oxidation of stannous chloride)

$2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3 + 2H_2O$

$Cu_2Cl_2 + Cl_2 \longrightarrow 2CuCl_2$

$2Sb + 3Cl_2 \longrightarrow 2SbCl_3$

(iv) **Removal of Electropositive element**

Example :

$2NaI + H_2O_2^* \longrightarrow 2NaOH + I_2$ (Oxidation of NaI)

$2KI + Cl_2^* \longrightarrow 2KCl + I_2$ (Oxidation of KI)

$2K_2MnO_4 + Cl_2^* \longrightarrow 2KCl + 2KMnO_4$

(Oxidation of Potassium-magnate)

(v) **Increment in oxidation state of Electropositive element**

Example :

$\underset{Sn^{+2}}{SnCl_2} + Cl_2^* \longrightarrow \underset{Sn^{+4}}{SnCl_4}$

Note : The substance which brings oxidation is know as oxidising agent. The substance marked with asterisk sign (*) in above equations are oxidising agents.

(b) **Reduction :** Reduction is a process which involves

(i) **Removal of Oxygen**

Example :

$CuO + C^* \longrightarrow Cu + CO$ (Reduction of cupric oxide)

$CuO + H_2 \longrightarrow Cu + H_2O$

$ZnO + H_2 \longrightarrow Zn + H_2O$

$\underset{\text{steam}}{H_2O} + \underset{\text{coke}}{C^*} \longrightarrow \underbrace{CO + H_2}_{\text{Water gas}}$ (Reduction of water)

$Fe_3O_4 + 4H_2^* \rightarrow 3Fe + 4H_2O$ (Reduction of Fe_3O_4)

(ii) **Addition of Hydrogen**

Example :

$N_2 + 3H_2 \longrightarrow 2NH_3$

$Cl_2 + H_2^* \longrightarrow 2HCl$ (Reduction of Chlorine)

$S + H_2^* \longrightarrow H_2S$ (Reduction of Sulphur)

(iii) **Removal of Electronegative element**

Example :

$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$

(Reduction of mercuric chloride)

$2FeCl_3 + H_2^* \longrightarrow 2FeCl_2 + 2HCl$

(Reduction of ferric chloride)

(iv) **Addition of Electropositive element**

$HgCl_2 + Hg^* \longrightarrow Hg_2Cl_2$

(Reduction of mercuric chloride)

$CuCl_2 + Cu^* \rightarrow Cu_2Cl_2$ (Reduction of cupric chloride)

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

(Reduction of mercuric chloride)

(v) **Decrement in oxidation state of Electropositive element**

Example :

$CuO + C^* \rightarrow Cu + CO$ (Reduction of cupric oxide)

$\underset{Cu^{+2}}{CuSO_4} + Fe^* \rightarrow FeSO_4 + \underset{Cu^0}{Cu}$ (Reduction of $CuSO_4$)

$\underset{Fe^{+3}}{FeCl_3} + H_2S^* \rightarrow \underset{Fe^{+2}}{FeCl_2} + 2HCl + S$ (Reduction of $FeCl_3$)

Note : The substance which brings reduction is known as reducing agent. The substances marked with asterisk sign (*) in the above equations are reducting agents.

2. **Modern concept or Electro concept :**

(i) **Oxidation** : According to this concept the process which involves the loss of one or more electrons from an atom or ion or molecule is called oxidation (de-electronation).

Example :

$Mg \longrightarrow Mg^{++} + 2e^-$

$\underset{\text{(ferrous)}}{Fe^{++}} \longrightarrow \underset{\text{(ferric)}}{Fe^{+++}} + e^-$

$2H_2O \longrightarrow 4H^{\oplus} + O_2 + 4e^-$

$2Cl^- \longrightarrow Cl_2 + 2e^-$

$[Fe(CN)_6]^{4-} \longrightarrow [Fe(CN)_6]^{3-} + e^-$

(ii) Reduction : According to this concept, the process which involves gain one or more electrons by an atom or ion or molecule is called reduction (electronation).

Example :

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

Stannic ion → Stannous ion

$2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

$ClO + e^- \longrightarrow Cl^-$

$MnO_4^- + e^- \longrightarrow MnO_4^{-2}$

$[Fe(CN)_6]^{3-} + 3e^- \longrightarrow [Fe(CN)_6]^{-6}$

$MnO_2 + 4H^{\oplus} + 2e^- \longrightarrow Mn^{+2} + 2H_2O$

3. Valency concept of oxidation reduction :

(i) Oxidation : According to this concept, increase in (+)ve valency or decrease in (–) ve valency in a reaction is called oxidation.

Example : $\underset{(0)}{2Mg} + O_2 \longrightarrow \underset{(+2)}{2MgO}$

(ii) Reduction : According to this concept, it is the process in which (+) ve valency decreases whereas (–)ve valency increases in a reaction is called reduction.

Example : $\underset{(+2)}{2HgCl_2} + SnCl_2 \longrightarrow \underset{(+1)}{Hg_2Cl_2} + SnCl_4$

4. Oxidation number concept :

(i) Oxidation : According to this concept, increase in oxidation no. in an element in a reaction is called oxidation.

(ii) Reduction : According to this concept, decrease in oxidation no. in an element in a reaction is called reduction.

Example : $\underset{+7}{KMnO_4} + \underset{+2}{FeSO_4} \longrightarrow \underset{+2}{MnO} + \underset{+3}{Fe_2(SO_4)_3}$

Note : Redox reaction involves two half reactions, one involving loss of electron or electrons called half oxidation reaction and the other involving gain of electron or electrons called half reduction reaction.

−2e$^-$
Oxidation

Example : $2Hg^{+2} + Sn^{+2} \longrightarrow Hg_2^{+2} + Sn^{+4}$

+2e$^-$
Reduction

OXIDISING AGENT OR OXIDANT (O.A.) :

It is the substance which accepts electrons in a chemical reaction i.e. electron acceptors are oxidising agent.

(i) Oxidising agents are lewis acids.

(ii) Substances which can oxidises others and reduces themselves.

(iii) Substances which shows the decrement in oxidation number.

Some Important oxidising agent or oxidant :

* All metallic oxides like Li_2O, Na_2O, Na_2O_2, CO_2, CaO, MgO, BaO_2 etc.
* Some nonmetallic oxides like CO_2, SO_2, H_2O_2, O_3.
* All neutral compounds or ions in which element shows their higher oxidation number or state act as oxidant or oxidising agent like $KMnO_4$, H_2SO_4, $SnCl_4$, H_3PO_4, $K_2Cr_2O_7$, $HClO_4$, $CuCl_2$, HNO_3, H_2SO_5, $FeCl_3$, $HgCl_2$, etc.
* Fluorine is the strongest oxidising substance.

Test of oxidising substance : When KI is added in an oxidising solution then the iodine is liberate and when dissolve in starch paper it gives blue colour.

REDUCTING AGENT OR REDUCTANT (R.A.)

The substance which donates electrons in a chemical reaction is called reducing agent i.e. electrons donors are reducing agents.

* Reducting agents are lewis bases.
* Substances which can reduce other and oxidises themselves.
* Substances which show the increment in oxidation number.

Some Important reducing agents or reductants :

* All metals like, K, Mg, Ca etc.
* All metallic hydrides like NaH, CaH_2, $LiAlH_4$, $NaBH_4$, AlH_3 etc.
* All hydroacids like HF, HCl, HBr, H_2S etc.
* Some organic compounds like Aldehyde, formic acid, oxalic acid, tartaric acid.
* All neutral compounds or ions, which show their lower oxidation state.
  MnO, HClO, $HClO_2$, H_3PO_2, HNO_2, H_2SO_3, $FeCl_3$, $SnCl_2$, Hg_2Cl_2, CH_2Cl_2 etc.
* In aqueous solution Li is the strongest reducing substance although in non aqueous solution Cs is the strongest reducing substance.

Test of reducing substance : When acidic $KMnO_4$ solution is added in reducing solution then the violet colour is disappeared or addition of acidic $K_2Cr_2O_7$ changes colour to green.

Note for the identification of oxidising and reducing agents

(i) If an element is in its highest possible oxidation state in a compound, the compound can function as an oxidising agent.
Example : $KMnO_4$, $K_2Cr_2O_7$, HNO_3, H_2SO_4, $HClO_4$ etc.

(ii) If an element is in its lowest possible oxidation state in a compound, the compound can function only as a reducing agent. Example: H_2S, $H_2C_2O_4$, $FeSO_4$, $Na_2S_2O_3$, $SnCl_2$ etc.

(iii) If an element is in its intermediate oxidation state in a compound, the compound can function both as an oxidising agent as well as reducing agent.
Example: H_2O_2, H_2SO_3, HNO_2, SO_2 etc.

(iv) If a highly electronegative element is in its highest oxidation state in a compound, that compound can function as a powerful oxidising agent.
Example: $KClO_4$, $KClO_3$, $KBrO_3$, KIO_3 etc.

(v) If an electronegative element is in its lowest possible oxidation state in a compound or in free state, it can function as a powerful reducing agent.
Example: I^-, Br^-, N^{3-} etc.

Some important compounds which can acts as oxidant and reductant both

HNO_2, SO_2, H_2O_2, O_3, Al_2O_3, CrO_2, MnO_2, ZnO, CuO

Note : Al_2O_3, CrO_2, MnO_2, ZnO, CuO are called as amphoteric oxides.

OXIDATION NUMBER

It represents the number of electrons gained or lost by atom when it changes in compound from a free state.

(i) If electrons are gained by an atom in the formation of compound, oxidation number is given (–) ve sign.

(ii) If electrons are lost by an atom in the formation of compound, oxidation number is given (+)ve sign.

(iii) It represents the real charge in case of ionic compounds and represents the imaginary charge in case of covalent compounds.

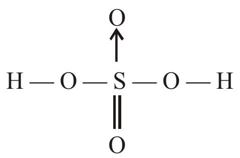
(iv) Maximum oxidation no. of an element is equal to group no. in the periodic table.

(v) Minimum oxidation no. of an element is equal to group no. – 8.

I Group elements always shows + 1 oxidation no.
II Group elements always shows + 2 oxidation no.
III Group elements always shows + 3 oxidation no.
IV Group shows –4 to +4 oxidation no.
V Group shows –3 to +5 oxidation no.
VI Group shows –2 to +6 oxidation no.
VII Group shows –1 to +7 oxidation no.
Inert gases are show zero oxidation no.

Oxidation no. for Coordinate bond :

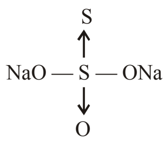
(i) When coordinate bond is formed from low electronegative element to high electronegative element then the e^- donor element shows +2 oxidation number whereas e^- acceptor element shows –2, oxidation no. in this type of bonded compounds. For example in H_2SO_4.



Here 'S' is low electronegative element than O therefore, oxidation number of S = +2 and O.N. of O = –2.

(ii) When coordinate bond is formed between the two same electronegative elements then the e^- donor element shows +2 oxidation number where e^-acceptor element shows –2 oxidation number in this type of bonded compound.

For example : In $Na_2S_2O_3$



Here O.N. of 'S' is +2, Because it is e^- donor and the other 'S' is –2, Because it is e^- acceptor.

(iii) When coordinate bond is formed from high electronegative element to low electronegative element then no change will be shown by both the elements, which are bonded by coordinate bond e.g. CH_3NC.

Oxidation State : Oxidation state of an atom is defined as oxidation number per atom for all practical purposes. Oxidation state is often expressed as oxidation number.

The rules to derive oxidation number or oxidation state :

(a) The O.S. of an element in its free state is zero. Example O.S. of Na, Cu, I, Cl, O etc. are zero

(b) Sum of O.S.of all the atoms in neutral molecule is zero.

(c) Sum of O.S. of all the atoms in a complex ion is equal number of charge present on it.

(d) In complex compounds, O.S. of some neutral molecules (ligands) is zero. Example CO, NO, NH_3, H_2O.

(e) Generally O.S. of Oxygen is –2 but in H_2O_2 it is –1.

(f) Generally O.S. of Hydrogen is +1 but in metallic hydrides it is –1.

(g) Generally O.S. of Halogen atoms is –1 but in interhalogen compounds it changes.

(h) Generally, O.S. of alkali metals is +1 and that of alkaline earth metals is +2.

(i) O.S. of transition elements vary from compound to compound. Mn has O.S. from +1 to +7.

$Mn_2O \rightarrow +1$, $MnO \rightarrow +2$, $Mn_3O_4 \rightarrow 8/3$,
$MnO_2 \rightarrow +4$, $Mn_2O_5 \rightarrow +5$, $MnO_4^{-2} \rightarrow +6$,
$MnO_4^- \rightarrow +7$

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

Oxidation state as a periodic property :

Oxidation state of an atom depends upon the electronic configuration of atom it is periodic properties.

(a) I A group or alkali metals shows +1 oxidation state.

(b) II A group or alkaline earth metals show +2 O.S.

(c) The maximum normal oxidation state, shown by III A group elements is +3. These elements also show +2 to +1 oxidation states.

(d) Elements of IVA group show their max and min. oxidation states +4 and –4 respectively.

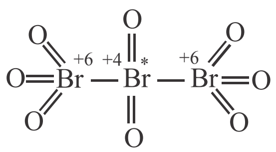
(e) Non metals shows number of oxidation states, the relation between max and min. oxidation states for non metals is equal to maximum O.S. – minimum O.S. = 8.
For example sulphur has maximum oxidation number +6 as being in VI A group element.

Fractional Oxidation states :

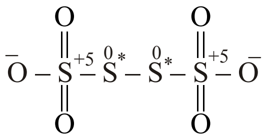
* Fractional oxidation state is the average oxidation state of the element under examination.
* Examples are:
  C_3O_2 : Oxidation number of carbon is (4/3)
  Br_3O_8 : oxidation number of bromine is (16/3)
  $Na_2S_4O_6$: oxidation number of sulphur is 2.5.
* The structural parameters reveal that the element for whom fractional oxidation state is realised is present in different oxidation states.
* Structure of the species C_3O_2, Br_3O_8 & $S_4O_6^{2-}$ reveal the following bonding situations:

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

Structure of C_3O_2 (carbon suboxide)



Structure of Br_3O_8 (tribromo octaoxide)



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

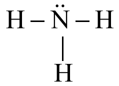
* In C_3O_2, two carbon atoms are present in +2 oxidation state each, whereas the third one is present in zero oxidation state and the average is 4/3. However, the realistic picture is +2 for two terminal carbons and zero for the middle carbon.

* 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. The average is 16/3.

* In the species $S_4O_6^{2-}$, each o f the two extreme sulphurs exhibits oxidation state of +5 and the two middle sulphurs as zero. The average of four oxidation numbers of sulphurs of the $S_4O_6^{2-}$ is 2.5, whereas the reality being + 5, 0, 0 and +5 oxidation number respectively for each sulphur.

Formal Charge : Formula for the formal charge of an atom in a compound : Formal charge = [Group number in periodic table or number of e⁻ i outer most orbital] – [Number of bonds + Number of unshared electrons]

e.g. formal charge on N in NH_3

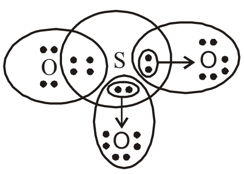


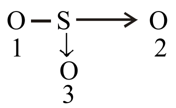
$= 5 - (3 + 2) = 0$

Example 1 :

Calculate the formal charge on each oxygen atom is SO_3.

Sol. Structure of SO_3 or Dot structure of SO_3 is





Formal charge on $O^1 = 6 - (2 + 4) = 0$

Formal charge on $O^2 = 6 - (1 + 6) = -1$

Formal charge on $O^3 = 6 - (1 + 6) = -1$

Formal charge on $S = 6 - (4 + 0) = +2$

Example 2 :

Find the oxidation number of cobalt in $[Co(NH_3)_6]\,Cl_2\,Br$.

Sol. Let the O.N. of Co be x

O.N. of NH_3 is zero ; O.N. of Cl is – 1 ; O.N. of Br is – 1

Hence, $x + 6(0) - 1 \times 2 - 1 = 0 \quad \therefore \quad x = +3$

so, the oxidation number of cobalt in the given complex compound is + 3.

Example 3 :

The order of increasing O.N. of S in S_8, $S_2O_8^{-2}$, $S_2O_3^{-2}$, $S_4O_6^{-2}$ is given below -

(A) $S_8 < S_2O_8^{-2} < S_2O_3^{-2} < S_4O_6^{-2}$

(B) $S_2O_8^{-2} < S_2O_3^{-2} < S_4O_6^{-2} < S_8$

(C) $S_2O_8^{-2} < S_8 < S_4O_6^{-2} < S_2O_3^{-2}$

(D) $S_8 < S_2O_3^{-2} < S_4O_6^{-2} < S_2O_8^{-2}$

Sol. **(D).** The O.Ns. of S are shown below along with the compounds

S_8	$S_2O_8^{-2}$,	$S_2O_3^{-2}$,	$S_4O_6^{-2}$
0	+7	+2	+2.5

Hence the order of increasing O.N. of S is -

$S_8 < S_2O_3^{-2} < S_4O_6^{-2} < S_2O_8^{-2}$

Example 4 :

Find the oxidation number of Cl in $NOClO_4$.

Sol. The compound may be written as $NO^+\,ClO_4^-$

For ClO_4^-, Let Ox. No. of Cl = a

$a + 4 \times (-2) = -1$

$a = +7$

Hence, the oxidation no. of Cl in $NOClO_4$ is + 7

Example 5 :

The two possible oxidation numbers of N atoms in NH_4NO_3 are respectively -

(A) +3, + 5 (B) +3, – 5

(C) – 3, + 5 (D) – 3, – 5

Sol. **(3).** There are two N atoms in NH_4NO_3, but one N atom has negative oxidation number (attached to H) and the other has positive oxid. no. (attached to O). Therefore evaluation should be made separately as -

O.N. of N is NH_4^+	Oxidn. no. of N in NO_3^-
$a + 4 \times (+1) = +1$	and $a + 3(-2) = -1$
$\therefore\ a = -3$	$\therefore\ a = +5$

Here the two O.N. are – 3 and + 5 respectively.

APPLICATIONS OF OXIDATION NUMBER

1. To compare the strength of acid and base :

Strength of acid $\propto$ Oxidation Number

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

Example 6 :

Find the order of acidic strength in HClO, $HClO_2$, $HClO_3$, $HClO_4$.

Sol.

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

$\therefore$ Strength of acid $\propto$ Oxidation Number

So the order will be

$HClO_4 > HClO_3 > HClO_2 > HClO$

2. To determine the oxidising and reducing nature :

Group	Range of Oxidation Number $(n-8)$ to n ; n = No. of Group
IA	$+1$
IIA	$+2$
IIIA	$+1, +3$
IVA	-4 to $+4$
VA	-3 to $+5$
VIA	-2 to $+6$ [Exception $\rightarrow$ Maximum Oxidation Number of Oxygen is 2]
VIIA	-1 to $+7$ [Exception $\rightarrow$ Oxidation Number of F is -1]

* If any compound is in maximum oxidation state (n), then it will act as only oxidant.
* If, compound is in minimum oxidation state $(n-8)$ then it will act as only reductant.
* If oxidation state is intermediate then compound can act as reductant as well as oxidant both.

For example,

(i) Oxidation No. of S in $H_2S = -2$ (min), So act as reductant.

(ii) Oxidation No. of N in $HNO_2 = +3$,
So act as reductant as well as oxidant both.
Oxidising strength $\propto$ Oxidation number

Reducing strength $\propto \dfrac{1}{\text{Oxidation number}}$

Common Oxidising and Reducing Agents :

Oxidising agent	**Effective Change**	**Decrease in O.N.**
$KMnO_4$ in acid sol^n	$MnO_4^- \rightarrow Mn^{2+}$	5
$KMnO_4$ in alkaline sol^n solution	$MnO_4^- \rightarrow MnO_2$	3
$K_2Cr_2O_7$ in acid sol^n	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$	3
Dilute HNO_3	$NO_3^- \rightarrow NO$	3
Concentrated HNO_3	$NO_3^- \rightarrow NO_2$	1
concentrated H_2SO_4	$SO_4^{2-} \rightarrow SO_2$	2
Manganese (IV) oxide	$MnO_2 \rightarrow Mn^{2+}$	2
Chlorine	$Cl \rightarrow Cl^-$	1
Chloric (I)acid	$ClO^- \rightarrow Cl^-$	2
KIO_3 in dilute acid	$IO_3^- \rightarrow I$	5
KIO_3 in concentrated acid	$IO_3^- \rightarrow I^-$	4

Redicing agent	**Effective Change**	**Increase in O.N.**
Iron(II) salts (acid)	$Fe^{2+} \rightarrow Fe^{3+}$	1
Tin (II) salts (acid)	$Sn^{2+} \rightarrow Sn^{4+}$	2
Sulphites (acid)	$SO_3^{2-} \rightarrow SO_4^{2-}$	2
Hydrogen sulphide	$S^{2-} \rightarrow S$	2
Iodidies (dilute acid)	$I^- \rightarrow I$	1
Iodides (concentrated acid)	$I^- \rightarrow I^+$	2
Metals, e.g. Zn	$Zn \rightarrow Zn^{2+}$	2
Hydrogen	$H \rightarrow H^+$	1

3. To determine the molecular formula of compound :

Example 7 :

Suppose that there are three atoms A, B, C and their oxidation numbers are 6, –1, –2, respectively then find the molecular formula of compound.

Sol. Since, the oxidation state of a free compound is zero.

$$x(6) + y(-1) + z(-2) = 0$$

Possible integers for x, y and z are

$$x = 1, y = 4, z = 1 \text{ and } x = 1, y = 2, z = 2$$

So molecular formula, AB_4C or AB_2C_2

REDOX REACTIONS

$\boxed{\text{RED}}$uction + $\boxed{\text{OX}}$idation

$\searrow \quad \swarrow$

REDOX

(a) The reactions in which oxidation and reduction both occur simultaneously are called redox reactions.

(b) Most of the chemical reactions are redox because if one element is there to lose electrons, other element has to be there to accept them.

(c) Any redox reaction may be divided in two parts.
(i) Oxidation half reaction (ii) Reduction half reaction

Some redox reaction.

S.No.	Reaction	Oxidant (Getting Reduced)	Reductant (Getting Oxidised)
1.	$C + O_2 \rightarrow CO_2$	O $[0 \rightarrow -2]$	C $[0 \rightarrow +4]$
2.	$PbS + 4O_3 \rightarrow PbSO_4 + 4O_2$	O $[+2 \rightarrow 0]$	S $[-2 \rightarrow +6]$
3.	$PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$	O $[-1 \rightarrow -2]$	S $[-2 \rightarrow +6]$
4.	$Sn + 2F_2 \rightarrow SnF_4$	F $[0 \rightarrow -1]$	Sn $[0 \rightarrow +4]$
5.	$SO_2 + 2H_2O + Cl_2 \rightarrow 2HCl + H_2SO_4$	Cl $[0 \rightarrow -1]$	S $[+4 \rightarrow +6]$
6.	$I_2 + 10HNO_3 \rightarrow 2HIO_3 + 10NO_2 + 4H_2O$	N $[+5 \rightarrow +4]$	I $[0 \rightarrow +5]$
7.	$CuO + H_2 \rightarrow Cu + H_2O$	Cu $[+2 \rightarrow 0]$	H $[0 \rightarrow +1]$
8.	$2KMnO_4 + 3H_2SO_4 + 5H_2S \rightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 5S$	Mn $[+7 \rightarrow +2]$	S $[-2 \rightarrow 0]$
9.	$H_2O_2 + Ag_2O \rightarrow 2Ag + H_2O + O_2$ (Oxygen of H_2O_2)	Ag $[+1 \rightarrow 0]$	O $[-1 \rightarrow 0]$
10.	$H_2SO_4 + 2HI \rightarrow SO_2 + I_2 + 2H_2O$	S $[+6 \rightarrow +4]$	I $[-1 \rightarrow 0]$

Note : In reaction 2 oxygens of ozone have different OS.
Structure of Ozone is

$O_3 \leftarrow O_1 = O_2$

OS of $O_1 = +2$

OS of $O_2 = 0$; Here O_1 is getting reduced in reaction 2

OS of $O_3 = -2$

Types of Redox reaction

(a) Combination redox reactions : The reactions in which two substances combine together to form a new compound are called combination reactions. For example, $A + B \rightarrow C$
For such a reaction to be redox reaction, either A or B, or both A and B should be in the elemental form.
For example, Burning of methane (or natural gas)

$$\overset{-4+1}{CH_4}(g) + \overset{0}{2O_2}(g) \rightarrow \overset{+4-2}{CO_2}(g) + \overset{+1\ -2}{2H_2O}(\ell)$$

Here carbon gets oxidised and oxygen gets reduced.

Reaction between magnesium and nitrogen:
Magnesium reacts with nitrogen at high temperature to give magnesium nitride.

$$\overset{0}{3Mg}(s) + \overset{0}{N_2}(g) \rightarrow \overset{+2\ -3}{Mg_3N_2}$$

Here magnesium gets oxidised and nitrogen gets reduced.

(b) Decomposition redox reactions : The reaction in which a compound breaks up into two or more substances at least one of which is an identical form is called a decomposition redox reaction. For example,

Decomposition of potassium chlorate :

$$\overset{+1\ +5\ -2}{2KClO_3}(s) \xrightarrow{\text{heat}} \overset{+1\ -1}{2KCl}(s) + \overset{0}{3O_2}(g)$$

Here, oxygen gets oxidised, and chlorine (in $KClO_3$) gets reduced.

Decomposition of sodium hydride :
Sodium hydride when heated gives off hydrogen gas.

$$\overset{+1\ -1}{2NaH}(s) \xrightarrow{\Delta} \overset{0}{2Na} + \overset{0}{H_2}(g)$$

In this reaction, sodium gets reduced, whereas hydrogen gets oxidised.

(c) Displacement redox reaction :
The reaction of the type

$AB + C \rightarrow AC + B$

in which an atom or ion in a compound is displaced by an ion (or atom) of another element, such that C and B are in elemental forms, is called displacement redox reaction.
There are two types of displacement redox reaction.

(i) Metal displacement reaction : A more electropositive metal can displace a less electropositive metal from its compound.
For example, **Reduction of Cr_2O_3 with aluminium :**

$$\overset{+3\ -2}{Cr_2O_3}(s) + 2Al(s) \xrightarrow{\Delta} \overset{+3\ -2}{Al_2O_3}(s) + \overset{0}{2Cr}(s)$$

This reaction forms the basis of the extraction of chromium by aluminium-reduction method.
The reduction of vanadium pentoxide (V_2O_5) to vanadium (V) by calcium, and that of $TiCl_4$ to Ti by magnesium are also displacement redox reactions.

(ii) Non-metal displacement reaction : There are many nonmetal displacement reactions.

Displacement of hydrogen :
Active metals displace hydrogen from water.
For example,

$$\overset{0}{2Na}(s) + \overset{+1\ -2}{2H_2O}(\ell) \longrightarrow \overset{+1\,-2+1}{2NaOH}(aq) + \overset{0}{H_2}(g)$$

$$\overset{0}{Ca}(s) + \overset{+1\ -2}{2H_2O}(\ell) \longrightarrow \overset{+2\ -1}{Ca(OH)_2}(aq) + \overset{0}{H_2}(g)$$

Magnesium & iron displace hydrogen from steam

$$\overset{0}{Mg}(s) + \underset{\text{steam}}{\overset{+1\ -2}{2H_2O}(\ell)} \longrightarrow \overset{+2\ \ -1}{Mg(OH)_2}(aq) + \overset{0}{H_2}(g)$$

$$\underset{\text{red hot}}{\overset{0}{3Fe}(s)} + \underset{\text{steam}}{\overset{+1\,-2}{4H_2O}(\ell)} \rightleftharpoons \overset{(+2,+3)-2}{Fe_3O_4}(s) + \overset{0}{4H_2}(g)$$

Certain metals such as Mg, Zn, Fe, etc. displace hydrogen from acids or acid solutions at room temperature.

$$\overset{0}{Mg}(s) + \overset{+1-1}{2HCl}(aq) \longrightarrow \overset{+2\ -1}{MgCl_2}(aq) + \overset{0}{H_2}(g)$$

$$\overset{0}{Zn}(s) + \overset{+1-1}{2HCl}(aq) \longrightarrow \overset{+2\,-1}{ZnCl_2}(aq) + \overset{0}{H_2}(g)$$

$$\overset{0}{Fe}(s) + \overset{+1-1}{2HCl}(aq) \longrightarrow \overset{+2\,-1}{FeCl_2}(aq) + \overset{0}{H_2}(g)$$

(d) Disproportionation reaction : When reduction and oxidation takes place on same element of a compound is called disproportionation reaction.
Example of disproportionation reactions are :

Decomposition of hydrogen perodixe :

$$\overset{+1\,-1}{2H_2O_2}(aq) \longrightarrow \overset{+1\,-2}{2H_2O}(\ell) + \overset{0}{O_2}(g)$$

In this reaction, the oxygen present in H_2O_2 (in –1 state) gets converted to elemental state (in zero oxidation state) and to –2 state in water molecule (H_2O).

$$\underset{O^{-1}}{H_2O_2} \longrightarrow \underset{O^{-2}}{H_2O} + \underset{O^{0}}{\tfrac{1}{2}O_2}$$

Reduction

Oxidation

Disproportionation redox reaction given by phosphorus, sulphur and chlorine :
Phosphorus, sulphur and chlorine undergo disproportionation in alkaline medium.

$$\overset{0}{P_4}(s) + 3OH^-(aq) + 3H_2O(\ell) \longrightarrow \overset{-3}{PH_3}(g) + 3H_2\overset{+1}{PO_2^-}(aq)$$

In this reaction, phosphorus (in elemental form) gas oxidised to +1 state and reduced to P^{3-} state.

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

In this reaction, the elemental sulphur gets oxidised to S^{2+} state and reduced to S^{2-} state.

$$\overset{0}{Cl_2}(s) + 2OH^-(aq) \longrightarrow \overset{+1}{ClO^-}(aq) + \overset{-1}{Cl^-}(aq) + H_2O(\ell)$$

In this reaction, the elemental chlorine gets oxidised to ClO^- (+1 state of Cl) and reduced to Cl^-, (–1 state of Cl). The hypochlorite ion (ClO^-) is a household bleaching agent.

Among the oxyanions of chlorine (ClO^-, ClO_2^-, ClO_3^-, ClO_4^-), ClO_4^- does not show disproportionation reaction. This is because in this ion, chlorine, occurs in its highest oxidation state (+7).

Fluorine does not exhibit any positive oxidation state. Therefore, it does not show disproportionation reaction.

Fluorine in alkaline solution gives the reaction,

$$2F_2(g) + 2OH^-(aq) \rightarrow 2F^-(aq) + OF_2(g) + H_2O(\ell)$$

(e) Comproportionation reaction : $NH_4NO_2 \rightarrow N_2 + 2H_2O$

Nitrogen in this compound has –3 and +3 oxidation number so it is not a definite value, so its not a disproportion reaction. It is an example of comproportionation reaction which is a case of redox reaction in which a element from two different oxidation state gets converted into a single oxidation state.

BALANCING OF REDOX EQUATION

There are two methods of balancing redox equations :

(1) Oxidation number change method.
(2) Ion electron method.

1. Oxidation number change method :

In a balanced redox reaction, total increase in oxidation number must be equal to decrease in oxidation number. This equivalence provides the basis for balancing redox reactions.

The general procedure 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 write 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 change in the reaction.

(v) In order to balance oxygen atoms, add H_2O molecules to the side deficient in oxygen. Then balance the number of H atoms by adding H^+ ions in the side deficient in hydrogen.

Example 8 :

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

Sol. Writing the oxidation number all the atoms.

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

there is change in oxidation number of Cu and N.

$$\overset{0}{Cu} + \longrightarrow \overset{+2}{Cu}(NO_3)_2 \quad \text{....... (1)}$$

(Oxidation no. is increment by 2)

$$H\overset{+5}{N}O_3 \longrightarrow \overset{+4}{N}O_2 \quad \text{......... (2)}$$

(Oxidation no. is decrement by 1)

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

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

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

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

This is the balanced equation.

2. Ion-Electron method :

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

(i) Writing the equation in ionic form.

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

(iii) Balance these half reactions separately.

(iv) Balance the atoms other than 'O' and 'H'.

(v) 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.

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

(vii) If the medium of reaction is basic, balance hydrogen atoms by adding H_2O to side having deficiency and an equal number of OH^- ions to the other side.

(viii) Balance the charge by electrons to the side which is rich in +ve charges. i.e. deficiency in electrons. Number of electrons added is equal to the deficiency.

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

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

Example :

(i) Acidic Medium :

(a) Consider the example,

$$Cr_2O_7^{2-} + C_2O_4^{2-} \xrightarrow{H^+} Cr^{3+} + CO_2$$

(b) Write both the half reaction.

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

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

(c) Atoms other than H and O are balanced

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

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

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

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

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

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

$Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$; $C_2O_4^{2-} \longrightarrow 2CO_2$

(f) Now, balance the charge by the addition of electrons (e).

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

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

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

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

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

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

(ii) Alkaline medium :

(a) Consider the reaction

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

(b) Separate the two half reactions.

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

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

(c) Balance O-atoms by adding H_2O.

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

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

(d) Balance H-atoms by adding H_2O to side having deficiency and an equal no. of OH^- ions to the other side ($\therefore$ medium is known)

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

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

(e) Balance the charges by electrons

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

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

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

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

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

$$\overline{4OH^- + 2Cr(OH)_3 + IO_3^- \rightarrow 5H_2O + 2CrO_4^{2-} + I^-}$$

Example 9 :

What will be the value of x, y and z in the following equation

$H_2C_2O_4 + xH_2O_2 \longrightarrow yCO_2 + zH_2O$

Sol. **(i)** The half reaction for oxidation is ,

$H_2C_2O_4 \longrightarrow CO_2$

Balancing carbon atoms on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2$

Balancing hydrogen atoms on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+$

Balancing the charge on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+ + 2e^-$ (balanced)

(ii) The half reaction for reduction is -

$H_2O_2 \longrightarrow H_2O$

Balancing oxygen atoms on both sides,

$H_2O_2 \longrightarrow 2H_2O$

Balancing hydrogen atoms, $H_2O_2 + 2H^+ \longrightarrow 2H_2O$

Balancing the charge,

$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$ (balanced)

Now, adding both equation,

$H_2C_2O_4 + H_2O_2 \longrightarrow 2CO_2 + 2H_2O$

This is balanced equation.

EQUIVALENT WEIGHT

Equivalent weight of a substance (element or compound) is defined as "The number of parts by weight of it, that will combine with or displace directly or indirectly 1.008 parts by weight of hydrogen, 8 parts by weight of oxygen, 35.5 parts by weight of chlorine or 108 part by weight of silver the equivalent parts by weight of another element".

$$\text{Equivalent weight of element} = \frac{\text{Atomic weight of element}}{\text{Valency of element}}$$

Equivalent weight of any element can variable

$$\text{Equivalent weight of acid} = \frac{\text{Molecular weight}}{\text{Basicity}}$$

$$\text{Equivalent weight of Base} = \frac{\text{Molecular weight}}{\text{Acidity}}$$

$$\text{Equivalent weight of Salt} = \frac{\text{Molecular weight}}{\text{Total charge on cation/anion}}$$

Equivalent weight of acidic salt

$$= \frac{\text{Molecular weight}}{\text{No. of H atom that can be substituted}}$$

Equivalent weight of compound

= Eq. wt. of cation + Eq. wt. of anion

Equivalent weights of oxidising and reducing agents.

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

(a) Equivalent wt. of an oxidant (get reduced)

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

$$= \frac{\text{Molecular weight}}{\text{Decrease in O.S.} \times \text{No. of atom undergoing reduction}}$$

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

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

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

Example 10 :

Find equivalent weight of $K_2Cr_2O_7$ in acidic medium

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

Sol. Here atoms undergoing reduction is Cr its O.S. is decreasing from 6 to 3

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

Note: [6 in denominator indicates that 6 electrons were gained by $Cr_2O_7^{2-}$ as it is clear from the given balanced equation]

Example 11:

Find equivalent weight of $KMnO_4$

(i) in acidic medium, (ii) in neutral medium,

(iii) in alkaline medium

Sol. (i) $\underset{(+7)}{MnO_4^-} \longrightarrow \underset{(+2)}{Mn^{2+}}$ (acidic medium)

Here 5 electrons are taken so eq. wt.

$$\frac{\text{Mol. wt. of } KMnO_4}{5} = \frac{158}{5} = 31.6$$

(ii) $\underset{(+7)}{MnO_4^-} \longrightarrow \underset{(+4)}{MnO_2}$ (neutral medium)

Here, only 3 electrons are gained, so eq. wt

$$= \frac{\text{Mol. wt. of } KMnO_4}{3} = \frac{158}{3} = 52.7$$

(iii) $\underset{(+7)}{MnO_4^-} \longrightarrow \underset{(+6)}{MnO_4^{-2}}$ (alkaline medium)

Here, only one electron is gained, so eq. wt

$$= \frac{\text{Mol. wt. of } KMnO_4}{1} = 158$$

Note : It is important to note that $KMnO_4$ acts as an oxidant in every medium although with different strength which follows the order –

acidic medium > neutral med. > alkaline medium

while, $K_2Cr_2O_7$ acts as an oxidant only in acidic medium as follows $Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+}$

$(2 \times 6) \longrightarrow (2 \times 3)$

Here, 6 electrons are gained so eq. wt

$$= \frac{\text{Mol. wt. of } K_2Cr_2O_7}{6} = \frac{294.21}{6} = 49.03$$

LAW OF EQUIVALENCE

The law states that one equivalent of an element combine with one equivalent of the other. In a chemical reaction equivalent and mill equivalent of reactants react in equal amount to give same no. equivalent or milli equivalent of products separately.

(i) $aA + bB \rightarrow mM + nN$

m.eq of A = m. eq. of B = m.eq of M = m.eq. of N

(ii) In a compound M_xN_y

m.eq of M_xN_y = m.eq of M = m.eq of N

Example 12 :

The number of moles of oxalate ions oxidized by one mole of MnO_4^- ion in acidic medium.

(A) 5/2 (B) 2/5

(C) 3/5 (D) 5/3

Sol. **(A).** Equivalents of $C_2O_4^{2-}$ = equivalents of MnO_4^-

x (mole) × 2 = 1 x 5

x = 5/2

OXIDATION-REDUCTION TITRATION (REDOX TITRATIONS)

* The chemical reactions proceed with transfer of electrons i.e, simultaneous loss of and gain of electrons among the reacting ions in solutions.
* **Equivalence point or End point :** The point at which the amounts of two reactants are just equivalent is called as end point of the titration.
* **Indicator :** The substance which helps in the detection of the end point or equivalence point of the titration is called as indicator. Those substances which undergo detectable change at the end point are used as indicators.
* Titrations according to reagent used –

(a) Potassium permanganate titration :

One of the most useful reactants for redox titrations is potassium permanganate, $KMnO_4$, especially when the reaction can be carried out in an acidic solution. Permanganate ion is a powerful oxidizing agent, so it oxidizes most substances that are capable of being oxidized. That's one reason why it is used. Especially important, though, is the fact that the MnO_4^- ion has a deep purple colour and its reduction product in acidic solution is the almost colourless Mn^{2+} ion. Therefore, when a solution of $KMnO_4$ is added from a buret to a solution of a reducing agent, the chemical reaction that occurs forms a nearly colourless product. As the $KMnO_4$ solution is added, the purple colour continues to be destroyed as long as there is any reducing agent left. However, after the last trace of the reducing agent has been consumed, the MnO_4^- ion in the next drop of titrant has nothing to react with, so it colours the solution pink. This signals the end of the titration. In this way, permanganate ion serves as its own indicator in redox titrations.

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

$2MnO_4^- + 16H^+ + 10Fe^{2+} \rightarrow 2Mn^{2+} + 2H_2O + 10Fe^{3+}$

(b) Potassium dichromate titration :

In place of potassium permanganate, potassium dichromate can also be used in the presence of dil. H_2SO_4. The ionic equation for the redox reaction with $FeSO_4$ (Fe^{2+} ions) is given. $C_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq)$

$\rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$

The reaction is used in the estimation of ferrous ions in volumetric analysis.

(c) Iodimetric and iodometric titrations :

The reduction of free iodine to iodide is done in these reactions, they are iodimetric titrations i.e, $I_2 + 2e^- \rightarrow 2I^-$

A solution of free Iodine is prepared by dissolving it in KI solution. This solution is standardised before use. This standard solution is then employed for estimation of sulphite, thiosulphate, arsenite etc. In iodometric titrations iodide is oxidised to iodine with an oxidising agent in acidic or neutral media. i.e., $2I^- - 2e^- \longrightarrow I_2$

The liberated iodine is then titrated with thiosulphate, oxyhalogens, dichromate cupric ion etc. During these titrations starch is used as an indicator. Starch solution gives blue or violet colour with iodine. At the end point the blue or violet colour disappears when the free iodine is converted to iodide.

(d) Sodium thiosulphate titration :

The redox reaction between sodium thiosulphate ($S_2O_3^{2-}$ ions) and I_2 are also an example of redox titration.

$$I_2\,(aq) + 2\,S_2O_3^{2-}\,(aq) \rightarrow 2I^-\,(aq) + S_4O_6^{2-}\,(aq)$$

This method based on the fact that iodine itself gives an intense blue colour with starch and has a very specific reaction with thiosulphate ($S_2O_3^{2-}$) ions.

ROLE OF REDOX REACTIONS IN HUMAN ACTIVITY

* A Breathalyzer test for alcohol makes use of the oxidation of alcohol by dichromate ion. The dichromate ion is orange, but its reduction product is green Cr^{3+} ion. The breath of a person who is intoxicated contains alcohol vapour, which passes through an acidic solution containing $Cr_2O_7^{2-}$ when the person blows through the Breathalyzer device. Any alcohol in the exhaled air is oxidized, which is signaled by the appearance of green Cr^{3+}. The more alcohol the person has consumed, the greater the intensity of the green colour.
* Metal oxides are reduced to metals by using suitable reducing agents. Fe_2O_3 is reduced to iron in a blast furnace using coke. Al_2O_3 is reduced to aluminium by cathodic reduction in an electrolytic cell.
* Combustion of a fuel is an oxidation reaction. Fuels on combustion (oxidation accompanied by heat and light) produce heat and light energies and thus are the most important source of energy that meets our daily needs.

  $$\text{Fuels (wood, gas, kerosene, petrol)} + O_2 \rightarrow CO_2 + H_2O + \text{Other products} + \text{Energy}$$

  In living cells, glucose, $C_6H_{12}O_6$ is oxidised to CO_2 and water in the presence of oxygen and energy is released. This energy is utilised by the body for doing physical and mental work.
* Many electrochemical cells based on redox reactions may be used for generating electricity. For example, electrical energy demand in a space capsule is met out of the reaction between hydrogen and oxygen used in fuel cells.

THE ACTIVITY SERIES

* List of metals arranged in order of decreasing ease of oxidation, is called an activity series.
* By observing what happens when samples of various metals are placed in contact with solutions of other metals, chemists have arranged the metals according to the relative ease or difficulty with which they can be oxidized in a single-displacement reaction.
* Metallic zinc reacts with iron salts, and metallic copper reacts with silver salts. Experimentally, it is found that zinc reacts with both copper salts & silver salts, producing Zn^{2+}.

  $$Zn(s) + Fe^{2+}\,(aq) \rightarrow Zn^{2+}\,(aq) + Fe\,(s)$$

  $$Cu(s) + 2Ag^+\,(aq) \rightarrow Cu^{2+}\,(aq) + 2Ag\,(s)$$
* Zinc therefore has a greater tendency to be oxidized than does iron, copper, or silver. Although zinc will not react with magnesium salts to give magnesium metal, magnesium metal will react with zinc salts to give zinc metal:

  $$Zn(s) + Mg^{2+}\,(aq) \not\rightarrow Zn^{2+}(aq) + Mg\,(s)$$

  $$Mg\,(s) + Zn^{2+}\,(aq) \rightarrow Mg^{2+}\,(aq) + Zn\,(s)$$
* Magnesium has a greater tendency to be oxidized than zinc does.
* Pairwise reactions of this sort are the basis of the activity series which lists metals and hydrogen in order of their relative tendency to be oxidized.
* When copper metal is placed in a solution of silver nitrate, a redox reaction forms silver metal and a blue solution of copper(II) nitrate. The oxidation of copper to copper ions is accompanied by the reduction of silver ions to silver metal. The silver metal is evident on the surface of the copper wire.
* Electron releasing tendency of the metals is in the order : Zn > Cu > Ag.
* The metals at the top of the series, which have the greatest tendency to lose electrons, are the alkali metals (group 1), the alkaline earth metals (group 2), and Al (group 13).
* In contrast, the metals at the bottom of the series, which have the lowest tendency to be oxidized, are the precious metals or coinage metals-platinum, gold, silver, and copper, and mercury, which are located in the lower right portion of the metals in the periodic table.
* **Table : Activity series of metals in aqueous solution**

	Metal	Oxidation Reaction	
React vigorously with cold H_2O to form H_2	Lithium	$Li\,(s) \rightarrow Li^+(aq) + e^-$	Ease of oxidation increases
	Potassium	$K\,(s) \rightarrow K^+(aq) + e^-$	
	Barium	$Ba\,(s) \rightarrow Ba^{2+}(aq) + 2e^-$	
	Calcium	$Ca\,(s) \rightarrow Ca^{2+}(aq) + 2e^-$	
	Sodium	$Na\,(s) \rightarrow Na^+(aq) + e^-$	
React with steam to form H_2	Magnesium	$Mg\,(s) \rightarrow Mg^{2+}(aq) + 2e^-$	
	Aluminum	$Al\,(s) \rightarrow Al^{3+}(aq) + 3e^-$	
	Manganese	$Mn\,(s) \rightarrow Mn^{2+}(aq) + 2e^-$	
	Zinc	$Zn\,(s) \rightarrow Zn^{2+}(aq) + 2e^-$	
	Chromium	$Cr\,(s) \rightarrow Cr^{3+}(aq) + 3e^-$	
	Iron	$Fe\,(s) \rightarrow Fe^{2+}(aq) + 2e^-$	
React with simple acids to form H_2	Cobalt	$Co\,(s) \rightarrow Co^{2+}(aq) + 2e^-$	
	Nickel	$Ni\,(s) \rightarrow Ni^{2+}(aq) + 2e^-$	
	Tin	$Sn\,(s) \rightarrow Sn^{2+}(aq) + 2e^-$	
	Lead	$Pb\,(s) \rightarrow Pb^{2+}(aq) + 2e^-$	
	Hydrogen	$H_2\,(s) \rightarrow 2\,H^+(aq) + 2e^-$	
Will not dissolve in simple acids	Copper	$Cu\,(s) \rightarrow Cu^{2+}(aq) + 2e^-$	
	Silver	$Ag\,(s) \rightarrow Ag^+\,(aq) + e^-$	
	Mercury	$Hg\,(l) \rightarrow Hg^{2+}(aq) + 2e^-$	
	Platinum	$Pt\,(s) \rightarrow Pt^{2+}\,(aq) + 2e^-$	
	Gold	$Au\,(s) \rightarrow Au^{3+}(aq) + 3e^-$	

* When using the activity series to predict the outcome of a reaction, keep in mind that any element will reduce compounds of the elements below it in the series. Because magnesium is above zinc in "The Activity Series", magnesium metal will reduce zinc salts but not vice versa.
* Similarly, the precious metals are at the bottom of the activity series, so virtually any other metal will reduce precious metal salts to the pure precious metals.

* Hydrogen is included in the series, and the tendency of a metal to react with an acid is indicated by its position relative to hydrogen in the activity series.
* Only those metals that lie above hydrogen in the activity series dissolve in acids to produce H_2.
For example, Ni reacts with HCl(aq) to form H_2:
$Ni(s) + 2HCl(aq) \rightarrow NiCl_2(aq) + H_2(g)$
Because elements below hydrogen in the activity series are not oxidized by H^+, Cu does not react with HCl (aq). Interestingly, copper does react with nitric acid, but the reaction is not oxidation of Cu by H^+ ions. Instead, the metal is oxidized to Cu^{2+} by the nitrate ion, accompanied by the formation of brown nitrogen dioxide, $NO_2(g)$:
$Cu(s)+4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(\ell) + 2NO_2(g)$
As the copper is oxidized in this reaction, NO_3^- where the oxidation number of nitrogen is +5, is reduced to NO_2, where the oxidation number of nitrogen is +4.
* The precious metals lie below hydrogen, they do not dissolve in dilute acid and therefore do not corrode readily.

Example 13 :
9.0 g of an ammonia solution is treated with 50 ml 0.5 N H_2SO_4 solution, 20 ml of 0.1 N NaOH is required for back titration. What is the percentage of ammonia in the solution?

Sol. meq. of acid taken initially = 50 × 0.5 = 25
meq. of NaOH used up in back titration = 20 × 0.1 = 2
meq. of H_2SO_4 reacted with NH_3 = 25.00 – 2.00 = 23
= meq of NH_3
$\Rightarrow$ Mass of ammonia = $23 \times 10^{-3} \times 17 = 0.391$ g
$\Rightarrow$ Mass percentage of ammonia = $\dfrac{0.391}{9} \times 100 = 4.34$

TRY IT YOURSELF

Q.1 In acting as a reducing agent, a piece of metal M weighing 16 grams gives 2.25×10^{23} electrons, what is the equivalent weight of the metal–
(A) 42.83 (B) 21.33
(C) 83.32 (D) 32

Q.2 What weight of HNO_3 is needed to convert 62gm of P_4 in H_3PO_4 in the reaction?
$P_4 + HNO_3 \rightarrow H_3PO_4 + NO_2 + H_2O$
(A) 63gm (B) 630gm
(C) 315gm (D) 126gm

Q.3 In the reaction: $C_2O_4^{-2} + MnO_4^- + H^+ \rightarrow Mn^{+2} + CO_2$ the reductant is –
(A) $C_2O_4^{-2}$ (B) H^+
(C) MnO_4^- (D) None of the above

Q.4 Oxidation number of cobalt in $[Co(NH_3)_6]Cl_2Br$ is –
(A) +6 (B) 0
(C) +3 (D) +2

Q.5 One mole of N_2H_4 loses 10 mole electrons to form a new compound Y. Assuming that all the N_2 appears in new compound, what is the oxidation state of Nitrogen in Y? (There is no change in the oxidation state of H)

Q.6 Number of moles of electrons produced per mole of $FeCr_2O_4$ in the following redox reaction.
$FeCr_2O_4 + KOH + O_2 \rightarrow K_2CrO_4 + Fe_2O_3$
(A) 6 (B) 1
(C) 7 (D) 3

Q.7 0.1 N $K_2C_2O_4.3H_2C_2O_4.4H_2O$ solution reacts completely with 20ml 0.05 M of $KMnO_4$ solution in acidic medium. Another sample of same solution of $K_2C_2O_4.3H_2C_2O_4.4H_2O$, having same volume is titrated with (1/8)M NaOH solution, then volume of NaOH solution is –
(A) 20ml (B) 30ml
(C) 50ml (D) None of these

Q.8 The number of moles of NaOH required to completely react with 20ml of 0.1 M NaH_2PO_3 solution and 40ml of 0.1M $NaHCO_3$.
(A) 8×10^{-3} (B) 2×10^{-3}
(C) 4×10^{-3} (D) 6×10^{-3}

Q.9 Which of the following acts as both oxidant and reductant
(A) HNO_3 (B) HNO_2
(C) Both HNO_3 & HNO_2 (D) Neither HNO_3 nor HNO_2

Q.10 State which of the following reactions is neither oxidation nor reduction
(A) $Na \rightarrow NaOH$ (B) $Cl_2 \rightarrow Cl^- + ClO_3^-$
(C) $P_2O_5 \rightarrow H_4P_2O_7$
(D) $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Q.11 The composition of a sample of wustite is $Fe_{0.93}O_{1.00}$. What percentage of iron is present in the form of Fe (III).
(A) 13.05 (B) 14.05
(C) 15.05 (D) 16.05

Q.12 $O_2 + FeS_2 \rightarrow FeO + SO_2$
The number of moles of FeS_2 required to produce 20 mole electrons to reduce O_2 in the above reaction is –
(A) 2 (B) 10
(C) 10/3 (D) 5/3

ANSWERS

(1) (A)	**(2)** (B)	**(3)** (A)
(4) (C)	**(5)** +3	**(6)** (C)
(7) (B)	**(8)** (D)	**(9)** (B)
(10) (C)	**(11)** (C)	**(12)** (A)

IMPORTANT POINTS

*

S.No.	Oxidation	Reduction
1.	Addition of oxygen.	Removal of oxygen.
2.	Removal of hydrogen.	Addition of hydrogen.
3.	Addition of an electronegative element.	Removal of an electronegative element.
4.	Removal of an electropositive element.	Addition of an electropositive element.
5.	Loss of electron.	Gain of electron.

* Oxidising property $\propto$ Tendency of accepting electron (s)
* Oxidising property $\propto$ Electronegativity
* Oxidising property $\propto$ Oxidation number

* Oxidising property $\propto \dfrac{1}{\text{Ionic or atomic size}}$
* Reducing property $\propto$ Tendency of losing electron(s)
* Reducing property $\propto$ Ionic or atomic size
* Reducing property $\propto \dfrac{1}{\text{Electronegativity}}$
* Reducing property $\propto \dfrac{1}{\text{Oxidation number}}$
* Decreasing order of reducing strength of metals
  $Zn > Fe > Pb > Cu > Ag$
* Decreasing order of oxidising strength of metal ion
  $Ag^{+} > Cu^{+2} > Pb^{+2} > Fe^{+2} > Zn^{+2}$
* Order of oxidising efficiency : $F_2 > Cl_2 > Br_2 > I_2$
* Order of reducing efficiency : $I^{-1} > Br^{-1} > Cl^{-1} > F^{-1}$
* When metals participate in chemical reactions, they are always oxidised. Thus metals always behave like reducing agents.
* The acidic nature of non-metal oxide increases with increasing O.S. of the non-metal.
* The strength of oxy-acid of an element increases with increasing oxidation state of the element.
* Strength of oxy-acids nitrogen increases in the order.

$H_2\overset{+1}{N_2}O_2 < H\overset{+3}{N}O_2 < H\overset{+5}{N}O_3$

Hyponitrous acid (weakest acid)   Nitrous acid   Nitric acid (strongest acid)

* An element in its lowest state in its compound behave like a reducing agent only.
  If the element is in its highest oxidation state, it behaves like an oxidising agent only.
* Metals exist in mixed oxidation states in non-stoichiometric compounds. For Example : O.S. of Fe in $Fe_{0.94}O$ is $+2$ & $+3$
* The oxidation state of iron in ferro compounds is +2 while in ferri compounds, it is in +3 oxidation state.
* When an element is in its highest oxidation state in a molecule or ion, it is a very good oxidising agent or electron acceptor, because it is possible to decrease the oxidation number of the element. Example, $Cr_2O_7^{2-}$, MnO_4^{-} etc.
* When an element is in its lowest oxidation state in a molecule or ion, it is a very good reducing agent or electron donor because it is possible to increase the oxidation number of the element. Example, H^{-}, Zn etc.
* Substances having elements in the intermediate oxidation states are capable of either gaining or losing electrons and so they can act both as oxidising and reducing agents. Examples, H_2O_2, O_3, SO_2 etc.
* When a strongly electronegative element is in its highest oxidation state in a chemical species, that species can act as one of the best oxidising agent. Examples: F_2, BrO_3^{-}, ClO_4^{-} etc.
* When a strongly electropositive element is in its lowest oxidation state in a chemical species, the species is capable of acting as one of the best reducing agents.
  Examples: NH_2OH, H_2O_2 $SnCl_2$, S^{2-} etc.
* The salts of 'ous' acids are good reducing agents.
  Examples: halites, sulphites, nitrites etc.
* The salts of 'ic' acids are good oxidising agents.
  Examples: halates, sulphates, nitrates etc.
* Salts of per acids are the most powerful oxidising agents.
  Examples : perchlorates, permanganate etc.
* The prefix hypo shows the lowest oxidation state of the central atom and so it acts as reducing agent.
  Example : H_3PO_2 (hypophosphorus acid).
* Metals always show positive oxidation states. For example, alkali and alkaline earth metals have a greater tendency to form positive ions due to their low ionisation potential.
* Fluorine, being most electronegative among all the elements, shows negative oxidation state only. It never shows positive oxidation state in any compound.
* To determine the oxidation numbers of elements in compounds, it should be noted that electrons are shifted to the atom of more electronegative element. For example, in PCl_5, phosphorus atom has an oxidation number of +5, while chlorine has –1, since the electronegativity of phosphorus (2.1) is less than that of chlorine (3.0).
  In Cl_2O, the chlorine atom has a lower electronegativity (3.0) than the oxygen atom (3.5). Thus chlorine atom in the compound Cl_2O will have an oxidation number of +1 and oxygen atom has –2.
* Atoms in their highest oxidation states can function as oxidising agents only. Examples are $KMnO_4$, $K_2Cr_2O_7$, K_2CrO_4, PbO_2 etc.
* Metal ions with highest oxidation numbers, such as Hg^{2+}, Fe^{3+},Cu^{2+} also act as oxidising agents.
* Nitric acid and concentrated sulphuric acid are oxidising agents.
* The order of acidic strength in HClO, $HClO_2$, $HClO_3$, $HClO_4$ is $HClO_4 > HClO_3 > HClO_2 > HClO$
* **Some nonzero oxidation states (numbers) :**

1	1	3	4	5	6	7	8	9	11	12
H	**H**	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Na**	**Mg**
+1 –1	+1 –1	+1	+2	+3	+4 +2 –4	+5 +4 +3 +2 +1 –3	–1 –2	–1	+1	+2

13	14	15	16	17	19	20	21	22	23	24
Al	**Si**	**P**	**S**	**Cl**	**K**	**Ca**	**Sc**	**Ti**	**V**	**Cr**
+3	+4 –4	+5 +3 –3	+6 +4 –2	+7 +5 +3 +1 –1	+1	+2	+3	+4 +3 +2	+5 +4 +3 +2	+6 +3 +2

25	26	27	28	29	30	31	32	33	34	35	36
Mn	**Fe**	**Co**	**Ni**	**Cu**	**Zn**	**Ga**	**Ge**	**As**	**Se**	**Br**	**Kr**
+7 +6 +4 +3 +2	+3 +2	+3 +2	+2	+2 +1	+2	+3	+4 –4	+5 +3 –3	+6 +4 –2	+7 +5 +3 +1 –1	+4 +2

ADDITIONAL EXAMPLES

Example 1 :

Which of the following acts as both oxidant and reductant

(A) HNO_3 (B) HNO_2

(C) Both HNO_3 & HNO_2 (D) Neither HNO_3 nor HNO_2

Sol. **(B).** O.N. of N in HNO_2 is + 3

Max. O.N. of N is + 5

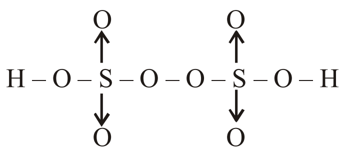
Min. O.N. of N is –3

Thus O.N. of N in HNO_2 can show an increase or decrease as the case may be. That is why HNO_2 acts as oxidant and reductant both. O.N. of N in HNO_3 is + 5. Hence it can act only as an oxidant.

Example 2 :

Find the oxidation number of S in $H_2S_2O_8$.

Sol. In $H_2S_2O_8$, two O atoms form peroxide linkage i.e.



$2 \times 1 + 2a + 6(-2) + 2(-1) = 0 \quad \therefore \; a = +6$

Thus the O.N. of S in $H_2S_2O_8$ is + 6

Example 3 :

The oxidation number of S in $(CH_3)_2$ SO is -

(A) 1 (B) 2

(C) 0 (D) 3

Sol. **(C).** Let the oxidation no. of S is 'a'

O.N. of $CH_3 = +1$; O.N. of $O = -2$

$2(+1) + a + (-2) = 0$; $a = 0$

Hence the oxidation no. of S in Dimethyl sulphoxide is zero.

Example 4 :

Balance the following equation by oxidation number method.

$$MnO_4^- + Fe^{+2} \longrightarrow Mn^{+2} + Fe^{+3}$$

Sol. Writing the oxidation number of all atoms.

$$\overset{+7\,-2}{MnO_4^-} + Fe^{+2} \longrightarrow Mn^{+2} + Fe^{+3}$$

Change in oxidation number has occurred in Mn and Fe.

$\overset{+7}{MnO_4^-} \longrightarrow Mn^{+2}$ .... (1) [Decrement in oxi. no. 5]

$Fe^{+2} \longrightarrow Fe^{+3}$ .... (2) [increment in oxi. no. by 1]

To make increase and decrease equal, equation (2) is multiplied by 5.

$$MnO_4^- + 5Fe^{+2} \longrightarrow Mn^{+2} + 5Fe^{+3}$$

To balance oxygen, $4H_2O$ are added to R.H.S. and to balance hydrogen, $8H^+$ are added to L.H.S.

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

This is the balanced equation.

Example 5 :

$KMnO_4$ solution is to be standardised by titration against As_2O_3 (s). A 0.1097g sample of As_2O_3 requires 26.10ml of the $KMnO_4$ solution for its titration. What are the molarity and normality of the $KMnO_4$ solution ?

Sol. $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$ (Reduction)

$As_2^{3+} \rightarrow 2As^{5+} + 4e^-$ (Oxidation)

$\therefore$ Meq of As_2O_3 = meq of $KMnO_4$

$$\frac{0.1097}{\frac{198}{4}} \times 1000 = 26.10 \times N \quad (E_{As_2O_3} = M/4)$$

$$\therefore N_{KMnO_4} = 0.085 \qquad \therefore N_{KMnO_4} = \frac{0.085}{5} = 0.017$$

Example 6 :

Which will be the value of x, y & z in the following equation-

$$xI_2 + yOH^- \longrightarrow IO_3^- + zI^- + 3H_2O$$

Sol.

$\overset{0}{I_2} \longrightarrow \overset{+5}{IO_3^-}$ .... (i) Oxidation

$\overset{0}{I_2} \longrightarrow \overset{1-}{I^-}$ .... (i) Reduction

Balancing atoms of Iodine on two sides, we get,

$I_2 \longrightarrow 2IO_3^-$

or $I_2 + 12OH \longrightarrow 2IO_3^- + 6H_2O$

Balancing change,

$I_2 + 12OH^- \longrightarrow 2IO_3^- + 6H_2O + 10e^-$

$I_2 \longrightarrow 2I^-$

and $(I_2 + 2e^- \longrightarrow 2I^-) \times 5$

Adding $6I_2 + 12OH^- \longrightarrow 2IO_3 + 10I^- + 6H_2O$

or, $3I_2 + 6OH^- \longrightarrow IO_3^- + 5I^- + 3H_2O$

Example 7 :

What will be the value of a, b, c, d, e and f in the following equation -

$$a\,MnO_4^- + bC_2O_4^{--} + cH^+ \longrightarrow dMn^{++} + eCO_2 + fH_2O$$

Sol. **(i)** The half reaction for reduction is, $MnO_4^- \longrightarrow Mn^{++}$

Balancing with respect to oxygen by adding $4H_2O$ on R.H.S., $MnO_4^- \longrightarrow Mn^{++} + 4H_2O$

Balancing with respect to hydrogen by adding $8H^+$ on L.H.S., $MnO_4^- + 8H^+ \longrightarrow Mn^{++} + 4H_2O$

Balancing charge by adding electrons,

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

(ii) The half reaction for oxidation is,

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

Balancing with respect to electrical charge by adding electrons on R.H.S.

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

Now, to equalise the number of electrons, the reduction half reaction is multiplied by 2 and oxidation half reaction by 5, so on adding, we get

$(C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-) \times 5$

$(MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{++} + 4H_2O) \times 2$

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

This is the balanced equation.

Example 8 :

9.8 ml of 0.1N $K_2Cr_2O_7$ + 2 ml of 0.01N $K_2Cr_2O_7$ solution is needed for the complete oxidation of 10 ml ferrous ammonium sulphate solution. What is the normality of Fe(II) ?

Sol. 2 ml of 0.01N $K_2Cr_2O_7 \equiv$ 2 ml. 0.01N $K_2Cr_2O_7$

$[Fe(II)\ N_1V_1 = N_2V_2\ [K_2Cr_2O_7]$

$N_1 \times 10 = 0.1N\,[9.8 + 0.2]$; $N_1 = 0.1N$

Example 9 :

q gm. of ferrous ammonium sulphate is present in one litre of the solution. 25 ml. of this solution is completely oxidised by 24 ml. of N/20 $KMnO_4$ solution. Supposing, ferrous ammonium sulphate is only 49% pure, the value of q would be –

(A) 20 (B) 40

(C) 10 (D) 60

Sol. **(B).** Normality of ferrous ammonium sulphate (FAS) is :

$N_1V_1 = N_2V_2$

$N_1 \times 25 = \frac{N}{20} \times 25$; $N_1 = \frac{N}{20}$

For FAS the reaction is

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

$\therefore$ Eq. Wt. of FAS = mol ey. go GSD = 392

$\because$ To get 49 gm. of FAS the sample reqd. = 100 g

$\therefore$ To get $\frac{392}{20}$ gm. of FAS the sample reqd.

$= \frac{100}{49} \times \frac{392}{20} = 40$ g

Example 10 :

Prove that following reactions are redox reactions.

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

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

(c) $4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3(s)$

(d) $2K(s) + F_2(g) \rightarrow 2K^+F^-(s)$

(e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

Sol. (a) $\overset{+2}{Cu}O(s) + \overset{0}{H_2}(g) \rightarrow \overset{0}{Cu}(s) + \overset{+1}{H_2}O$

Oxidation; Reduction

(b) $\overset{+3}{Fe_2}O_3(s) + 3\overset{+2}{C}O(g) \rightarrow 2\overset{0}{Fe}(s) + 3\overset{+4}{C}O_2$

Oxidation; Reduction

(c) $4BCl_3(g) + 3\ LiAlH_4(s) \rightarrow 2B_2H_6 + 3LiCl(s) + 3AlCl_3(s)$

Reduction (Addition of H); Removal of hydrogen (oxidation)

(d) $2K(s) + F_2(g) \rightarrow 2K^+F^-(s)$

Oxidation; Reduction

(e) $4\overset{-3}{N}H_3(g) + 5\overset{0}{O_2}(g) \rightarrow 4\overset{+2}{N}\overset{-2}{O}(g) + 6H_2O(g)$

Reduction; Oxidation

Example 11

Hydrogen peroxide in its reaction with KIO_4 and NH_2OH respectively, is acting as a –

(A) Reducing agent, oxidising agent

(B) Reducing agent, reducing agent

(C) Oxidising agent, oxidising agent

(D) Oxidising agent, reducing agent

Sol. **(A).** $KIO_4 + H_2O_2 \rightarrow KIO_3 + H_2O + O_2$

$2NH_2OH + H_2O_2 \rightarrow N_2 + 4H_2O$

H_2O_2 acting as a reducing agent with KIO_4 and H_2O_2 acting as an oxidising agent with NH_2OH.

Example 12

For the reaction :

$I^- + ClO_3^- + H_2SO_4 \longrightarrow Cl^- + HSO_4^- + I_2$

The correct statements(s) in the balanced equation is/are

(A) Stoichiometric coefficient of HSO_4^- is 6

(B) Iodide is oxidized

(C) Sulphur is reduced

(D) H_2O is one of the products

Sol. **(ABD).**

Balanced chemical equation is

$6I^- + ClO_3^- + 6H_2SO_4 \longrightarrow Cl^- + 6HSO_4^- + 3I_2 + 3H_2O$

Here, H_2O is produced and I^- is oxidized.

Example 13 :

The oxidation numbers-of sulphur in S_8, SO_2 and H_2S, respectively are

(A) 0, + 6, –2 (B) 0, + 4, –2

(C) 0, +1, +2 (D) 0, +1, –2

Sol. **(B).** The oxidation state of an element in its free state is zero, i.e., oxidation state of S in S_8 is 0.

Let the oxidation state of S in SO_2 and H_2O is x.

SO_2	H_2S
x+(−2)×2=0	(1)×2+x=0
x−4=0	2+x=0
x=+4	x=−2

Thus, the oxidation states of S in S_8 , SO_2 and H_2S are respectively 0, + 4 and –2.

QUESTION BANK

CHAPTER 8 : REDOX REACTIONS

EXERCISE - 1 [LEVEL-1]

Choose one correct response for each question.

PART 1 : BASIC OF REDOX REACTIONS

Q.1 Oxidation is –
(A) loss of electrons
(B) addition of hydrogen
(C) decrease in oxidation number
(D) all of the above

Q.2 Which of the following reactions takes place at anode?
(A) Reduction (B) Oxidation
(C) Decomposition (D) Dissolution

Q.3 Reduction involves –
(A) addition of oxygen (B) removal of hydrogen
(C) Both (A) and (B) (D) None of the above

Q.4 Oxidation is defined as –
(A) addition of electropositive element.
(B) addition of electronegative element.
(C) removal of electropositive element.
(D) Both (B) and (C).

Q.5 In the reaction: $Cl_2 + OH^- \rightarrow Cl^- + ClO_4^- + H_2O$
(A) Chlorine is oxidised.
(B) Chlorine is reduced.
(C) Chlorine is oxidised as well as reduced.
(D) Chlorine is neither oxidised nor reduced.

Q.6 Choose the correct statement –
(A) N_2 (g) + O_2 (g) $\rightarrow$ 2 NO (g) is an example of combination redox reactions.
(B) NaH(s) + $H_2O(l)$ $\rightarrow$ NaOH(aq) + H_2 (g) is an example of displacement redox reaction.
(C) $2NO_2(g) + 2OH^-(aq) \rightarrow NO_2^-(aq) + NO_3^-(aq) + H_2O(l)$ is an example of disproportionation redox reaction.
(D) All of these

Q.7 In the given reaction, $2Na + S \rightarrow Na_2S$, sulphur is
(A) oxidised (B) reduced
(C) oxidising agent (D) None of the above

Q.8 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) $2NO_2 + 2OH^- \rightarrow NO_2^- + NO_3^- + H_2O$

Q.9 In the reaction $3CuO + 2NH_3 \rightarrow N_2 + 3H_2O + 3Cu$ the change of NH_3 to N_2 involves
(A) Loss of 6 electrons per mol of N_2.
(B) Loss of 3 electrons per mol of N_2.
(C) Gain of 6 electrons per mol of N_2.
(D) Gain of 3 electrons per mol of N_2.

Q.10 The reaction, $2Na + Cl_2 \rightarrow 2Na^+Cl^-$ involves
(A) loss of $2e^-$ between $2Na \rightarrow 2Na^+$
(B) gain of $2e^-$ between $Cl_2 \rightarrow 2Cl^-$
(C) gain of $2e^-$ between $2Na \rightarrow 2Na^+$
(D) Both (A) and (B)

Q.11 When Sn^{2+} changes to Sn^{4+} in a reaction –
(A) It loses two electrons (B) It gains two electrons
(C) It loses two protons (D) It gains two protons

Q.12 In acid solution, the reaction $MnO_4^- \rightarrow Mn^{2+}$ involves
(A) Oxidation by 3 electrons
(B) Reduction by 3 electrons
(C) Oxidation by 5 electrons
(D) Reduction by 5 electrons

Q.13 When a strip of metallic zinc is placed in an aqueous solution of copper nitrate, the ions formed are of –
(A) Zn^{2+} (B) Cu^{2+}
(C) NO_3^- (D) Both (A) and (B)

Q.14 Which of the following is a redox reaction ?
(A) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$
(B) $CaC_2O_4 + 2\,HCl \rightarrow CaCl_2 + H_2C_2O_4$
(C) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + NHOH$
(D) $Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$

Q.15 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) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

Q.16 In reaction, $4Na + O_2 \rightarrow 2Na_2O$, sodium is oxidised, which acts as a
(A) oxidising agent (B) reducing agent
(C) Both (A) and (B) (D) All of these

Q.17 When a piece of sodium metal is dropped in water, hydrogen gas evolved because
(A) Sodium is reduced and acts as an oxidising agent.
(B) Water is oxidised and acts as a reducing agent.
(C) Sodium loses electrons and is oxidised while water in reduced.
(D) Water loses electrons and is oxidised to hydrogen.

PART - 2 : OXIDATION NUMBER

Q.18 The oxidation number and valency of C in sucrose ($C_{12}H_{22}O_{11}$) is –
(A) + 4 and 4 (B) +3 and 0
(C) +2 and 4 (D) 0 and 4

Q.19 Which of the following statements is correct regarding redox reactions?
(A) An increase in oxidation number of an element is called reduction.
(B) A decrease in oxidation number of an element is called oxidation.
(C) A reagent which lowers the oxidation number of an element in a given substance is reductant.
(D) A reagent which increases the oxidation number of an element in a given substance is reductant.

Q.20 The oxidation number of oxygen is KO_3, Na_2O_2 is –
(A) 3, 2 (B) 1, 0
(C) 0, 1 (D) –0.33, –1

Q.21 Consider the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
Choose the correct statement.
(A) Zn is reduced to Zn^{2+} (B) Zn is oxidised to Zn^{2+}
(C) Zn^{2+} is oxidised to Zn (D) Cu^{2+} is oxidised to Cu

Q.22 In the chemical reaction,
$Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$
(A) water is oxidised (B) silver is oxidised
(C) silver is reduced (D) hydrogen is reduced

Q.23 Oxidation number of P in KH_2PO_2 is
(A) +1 (B) +3
(C) +5 (D) –4

Q.24 The oxidation number of S in $(CH_3)_2$ SO is –
(A) 1 (B) 2
(C) 0 (D) 3

Q.25 What is the oxidation state of nitrogen in NaN_3?
(A) – 3/1 (B) 3
(C) – 3 (D) – 1/3

Q.26 What is the oxidation number of oxygen in OF_2?
(A) +2 (B) +4
(D) +3 (D) None

Q.27 M^{+3} ion loses $3e^-$. Its oxidation number will be
(A) 0 (B) + 3
(C) + 6 (D) – 3

Q.28 The oxidation number of S in $H_2S_2O_8$ is –
(A) +8 (B) –8
(C) +6 (D) +4

Q.29 Oxidation state of oxygen in F_2O is
(A) + 1 (B) + 2
(C) –1 (D) –2

Q.30 The oxidation number of nitrogen in NH_2OH is
(A) + 1 (B) – 1
(C) – 3 (D) – 2

Q.31 The oxidation number of C in $CH_3 – Cl$ is
(A) – 3 (B) – 2
(C) – 1 (D) 0

Q.32 Oxidation number of S in S_2Cl_2 is
(A) + 1 (B) – 1
(C) + 6 (D) 0

Q.33 Oxidation state of Fe in Fe_3O_4 is
(A) 3/2 (B) 4/5
(C) 5/4 (D) 8/3

Q.34 The oxidation number of nickel in $K_4[Ni(CN)_4]$ is –
(A) – 2 (B) – 1
(C) + 2 (D) 0

Q.35 Which of the following elements never show positive oxidation number
(A) O (B) Fe
(C) Ga (D) F

Q.36 The oxidation state of I in IPO_4 is
(A) +1 (B) +3
(C) +5 (D) +7

Q.37 The oxidation state of nitrogen in N_3H is –
(A) –3 (B) +3
(C) –1 (D) –1/3

Q.38 Identify the element that exhibits only –ve oxidation state.
(A) Cs (B) Ne
(C) I (D) F

Q.39 A compound contains atoms X, Y and Z. The oxidation number of X is +2, Y is +5 and Z is –2. The possible formula of the compound is
(A) XYZ_2 (B) $Y_2(XZ_3)_2$
(C) $X_3(YZ_4)_2$ (D) $X_3(Y_4Z)_2$

Q.40 Which of the following is not a redox reaction ?
(A) $CaCO_3 \rightarrow CaO + CO_2$
(B) $O_2 + 2H_2 \rightarrow 2H_2O$
(C) $Na + H_2O \rightarrow NaOH + \frac{1}{2}H_2$
(D) $MnCl_3 \rightarrow MnCl_2 + \frac{1}{2}Cl_2$

Q.41 In the reaction $PCl_3 + Cl_2 \rightarrow PCl_5$
(A) PCl_3 is acting as reductant.
(B) Cl_2 is acting as reductant.
(C) Both PCl_3 and Cl_2 are acting as reductant.
(D) Both PCl_3 and Cl_2 are acting as oxidant.

Q.42 The oxidation state of nitrogen in NO_3^- is
(A) –1 (B) +2
(C) +3 (D) +5

Q.43 If an atom is reduced, its oxidation number
(A) does not change (B) increases
(C) sharply decreases (D) slightly decreases

Q.44 The oxidation number of carbon in CH_2O is
(A) – 2 (B) + 2
(C) 0 (D) + 4

Q.45 The oxidation number of Mn in $MnSO_4$ is
(A) +5 (B) +7
(C) +4 (D) +2

Q.46 Carbon has oxidation number of __ and ___ respectively in diamond and graphite.
(A) +2, +2 (B) +1, +1
(C) +4, – 4 (D) zero, zero

Q.47 What is the oxidation number of chlorine in ClO_3^-
(A) +5 (B) +3
(C) +4 (D) +2

Q.48 Oxidation number of S in S^{2-} is
(A) –2 (B) 0
(C) –6 (D) +2

Q.49 Oxidation number of Xe in Ba_2XeO_6 is
(A) + 8 (B) + 10
(C) + 4 (D) + 3

Q.50 What is oxidation number of Fe in $Fe(CO)_5$?
(A) Zero (B) 5
(C) –5 (D) +3

PART - 3 : BALANCING REDOX REACTIONS

Q.51 The set of numerical coefficients that balances the equation: $K_2CrO_4 + HCl \rightarrow K_2Cr_2O_7 + KCl + H_2O$
(A) 1, 1, 2, 2, 1
(B) 2, 2, 1, 1, 1
(C) 2, 1, 1, 2, 1
(D) 2, 2, 1, 2, 1

Q.52 $C_2H_6(g) + nO_2 \rightarrow CO_2(g) + H_2O(\ell)$
In this equation, ratio of the coefficients of CO_2 and H_2O is
(A) 1 : 1
(B) 2 : 3
(C) 3 : 2
(D) 1 : 3

Q.53 What will be the value of x, y and z in the following equation: $xI_2 + yOH^- \rightarrow IO_3^- + zI^- + 3H_2O$
(A) 3, 5, 6
(B) 5, 6, 3
(C) 3, 6, 5
(D) 6, 3, 5

Q.54 In a balanced equation
$H_2SO_4 + x\,HI \rightarrow H_2S + y\,I_2 + z\,H_2O$, the values of x, y, z
(A) $x = 3, y = 5, z = 2$
(B) $x = 4, y = 8, z = 5$
(C) $x = 8, y = 4, z = 4$
(D) $x = 5, y = 3, z = 4$

Q.55 Which of the following is correctly balanced half reaction–
(A) $AsO_3^{-3} + H_2O \rightarrow AsO_4^{-3} + 2H^+ - 2e^-$
(B) $H_2O_2 + 2e \rightarrow O_2 + 2H^+$
(C) $Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O - 6e^-$
(D) $IO^-_3 + 6H^+ \rightarrow I_2 + 3H_2O + 5e^-$

Q.56 $xMnO_4^- + yH_2SO_4 \rightarrow 2Mn^{2+} + 5H_2O + 9O_2 + ze^-$
In this reaction values of x, y and z are
(A) 2, 5, 6
(B) 5, 2, 9
(C) 3, 5, 5
(D) 2, 6, 6

Q.57 The reaction is balanced if,
$5H_2O_2 + XClO_2 + 2OH^- \rightarrow XCl^- + YO_2 + 6H_2O$
(A) X = 5, Y = 2
(B) X = 2, Y = 5
(C) X = 4, Y = 10
(D) X = 5, Y = 5

Q.58 To balance the reaction, $Na + O_2 \rightarrow Na_2O$ the number of moles of sodium and disodium oxide are respectively–
(A) 2 and 2
(B) 4 and 2
(C) 2 and 4
(D) 3 and 5

Q.59 When copper is treated with a certain concentration of nitric acid, nitric oxide and nitrogen dioxide are liberated in equal volumes according to the equation,
$x\,Cu + y\,HNO_3 \rightarrow xCu(NO_3)_2 + NO + NO_2 + 3H_2O$
The coefficients x and y are
(A) 2 and 3
(B) 2 and 6
(C) 1 and 3
(D) 3 and 8

Q.60 In the ionic equation, $BiO_3^- + 6H^+ + xe^- \rightarrow Bi^{3+} + 3H_2O$ the value of x is
(A) 6
(B) 2
(C) 4
(D) 3

PART - 4 : REDOX REACTIONS AND ELECTRODE PROCESSES

Q.61 Which of the following is the strongest oxidizing agent
(A) F_2
(B) Cl_2
(C) Br_2
(D) I_2

Q.62 Negative $E^\ominus$ indicates that redox couple is –
(A) Weaker reducing agent than H^+/H_2 couple.
(B) Stronger reducing agent than H^+/H_2 couple.
(C) Stronger oxidising agent than H^+/H_2 couple.
(D) Weaker oxidising agent than H^+/H_2 couple.

Q.63 The strongest reducing agent is –
(A) HNO_2
(B) H_2S
(C) H_2SO_4
(D) $SnCl_2$

Q.64 Given $E^\circ_{Ag^+/Ag} = +0.80V$,
$E^\circ_{Cu^{2+}/Cu} = +0.34V$, $E^\circ_{Fe^{3+}/Fe^{2+}} = +0.76V$,
$E^\circ_{Ce^{4+}/Ce^{3+}} = +1.60V$
Which of the following statements is **not correct**?
(A) Fe^{3+} does not oxidise Ce^{3+}.
(B) Cu reduces Ag^+ to Ag.
(C) Ag will reduce Cu^{2+} to Cu.
(D) Fe^{3+} reduces Cu^{2+} to Cu.

Q.65 Al_2O_3 is reduced to aluminium by cathodic reduction in an electrolytic cell. Which of the following is NOT TRUE about this process?
(A) It involves a redox reaction.
(B) It involves the application of redox reactions in the extraction of metals.
(C) Here Al^{3+} is reduced to Al metal
(D) Here Al is oxidised to Al^{3+} ions

Q.66 E° values of some redox couples are given below. On the basis of these values choose the correct option. E° 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

Q.67 Given that $E^\circ_{K^+/K} = -2.93V$,
$E^\circ_{Fe^{2+}/Fe} = -0.44V$, $E^\circ_{Zn^{2+}/Zn} = -0.76\,V$
$E^\circ_{Cu^{2+}/Cu} = 0.34V$. Based on this data, which of the following is the strongest reducing agent?
(A) Cu (s)
(B) K^+(aq)
(C) Zn^{2+} (aq)
(D) Fe (s)

Q.68 Which of these substances is a good reducing agent?
(A) HI
(B) KBr
(C) $FeCl_3$
(D) $KClO_3$

EXERCISE - 2 [LEVEL-2]

Choose one correct response for each question.

Q.1 When $K_2Cr_2O_7$ is converted into K_2CrO_4 the change in oxidation number of Cr is –
(A) 0 (B) 3
(C) 4 (D) 6

Q.2 In which of the following reactions, the underlined substance has been reduced?
(A) $\underline{CO} + CuO \rightarrow \underline{CO_2} + Cu$
(B) $\underline{Cu}O + 2HCl \rightarrow \underline{Cu}Cl_2 + H_2O$
(C) $4\underline{H_2}O(g) + 3Fe \rightarrow 4\underline{H_2}(g) + Fe_3O_4$
(D) $\underline{C} + 4HNO_3 \rightarrow \underline{C}O_2 + 2H_2O + 4NO_2$

Q.3 Which of the following acts as both oxidant and reductant–
(A) HNO_3 (B) HNO_2
(C) Both (A) and (B) (D) None of these

Q.4 State which of the following reactions is neither oxidation nor reduction –
(A) $Na \rightarrow NaOH$
(B) $Cl_2 \rightarrow Cl^- + ClO^-_3$
(C) $P_2O_5 \rightarrow H_4P_2O_7$
(D) $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Q.5 The compound $YBa_2Cu_3O_7$ which shows super conductivity has copper in oxidation state ..... Assume that the rare earth element Yttrium is in its usual +3 oxidation state
(A) 3/7 (B) 7/3
(C) 3 (D) 7

Q.6 Arrange the following metals in which they displace each other from the solutions of their salts in decreasing order. Al, Cu, Fe, Mg & Zn.
$[E^\circ_{Al^{3+}/Al} = -1.66V,\ E^\circ_{Cu^{2+}/Cu} = +0.34V,$
$E^\circ_{Fe^{2+}/Fe} = -0.44V,\ E^\circ_{Mg^{2+}/Mg} = -2.36V$
and $E^\circ_{Zn^{2+}/Zn} = -0.76V\ ]$
(A) Cu, Fe, Zn, Al, Mg (B) Fe, Zn, Cu, Al, Mg
(C) Mg, Cu, Fe, Zn, Al (D) Mg, Al, Zn, Fe, Cu

Q.7 In the reaction $Al + Fe_3O_4 \rightarrow Al_2O_3 + Fe$ – what is the total no. of electrons transferred during the change –
(A) 16 (B) 24
(C) 8 (D) 12

Q.8 In the redox reaction :
$$10FeC_2O_4 + x\,KMnO_4 + 24H_2SO_4 \rightarrow 5Fe_2(SO_4)_3 + 20CO_2 + y\,MnSO_4 + 3K_2SO_4 + 24H_2O.$$
The values of x and y are respectively –
(A) 6, 3 (B) 3, 6
(C) 3, 3 (D) 6, 6

Q.9 A solution containing 2.68 x 10^{-3} mol of A^{+n} ions requires 1.61 x 10^{-3} mole of MnO^-_4 for the oxidation of A^{+n} to AO^-_3 in acidic medium. What is the value of n
(A) 1 (B) 2
(C) 3 (D) 4

Q.10 Which one of the following reactions does not involve either oxidation or reduction –
(A) $VO_2^+ \rightarrow V_2O_3$ (B) $Na \rightarrow Na^+$
(C) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$ (D) $Zn^{2+} \rightarrow Zn$

Q.11 Of the four oxyacids of chlorine the strongest oxidising agent in dilute aqueous solution is
(A) $HClO_4$ (B) $HClO_3$
(C) $HClO_2$ (D) HOCl

Q.12 A solution of sulphur dioxide in water reacts with H_2S precipitating sulphur. Here sulphur dioxide acts as
(A) As oxidising agent (B) A reducing agent
(C) An acid (D) A catalyst

Q.13 Oxidation number of O in H_2O_2 will be –
(A) –2 (B) –1
(C) +1 (D) +2

Q.14 Which is the best description of the behaviour of bromine in the reaction : $H_2O + Br_2 \rightarrow HOBr + HBr$
(A) Oxidised only
(B) Reduced only
(C) Proton acceptor only
(D) Both oxidised and reduced

Q.15 The valency of Cr in the complex $[Cr(H_2O)_4Cl_2]^+$
(A) 1 (B) 3
(C) 5 (D) 6

Q.16 Oxidation number of N in $(NH_4)_2SO_4$ is –
(A) –1/3 (B) –1
(C) +1 (D) –3

Q.17 The oxidation number of Mn in MnO_4^{-1} is
(A) + 7 (B) – 5
(C) + 6 (D) + 5

Q.18 The oxidation number of Pt in $[Pt(C_2H_4)Cl_3]^-$ is
(A) + 1 (B) + 2
(C) + 3 (D) + 4

Q.19 Which of the following is the strongest oxidising agent
(A) $BrO_3^- / Br^{2+}, E^o = +1.50$
(B) $Fe^{3+} / Fe^{2+}, E^o = +0.76$
(C) $MnO_4^- / Mn^{2+}, E^o = +1.52$
(D) $Cr_2O_7^{2-} / Cr^{3+}, E^o = +1.33$

Q.20 Which substance is serving as a reducing agent in the following reaction?
$14H^+ + Cr_2O_7^{2-} + 3Ni \rightarrow 2Cr^{3+} + 7H_2O + 3Ni^{2+}$
(A) H_2O (B) Ni
(C) H^+ (D) $Cr_2O_7^{2-}$

Q.21 Select the correct coefficients in the given reaction, when it is balanced.
$$xMnO_4^- + yH_2O_2 + z\,H^+ \rightarrow pMn^{2+} + qO_2 + rH_2O$$
(A) x = 1, y = 5, z = 8 (B) x = 5, y = 1, z = 8
(C) x = 1, y = 5, z = 6 (D) x = 2, y = 5, z = 6

Q.22 In the reaction, $HAsO_4 + Sn^{2+} \rightarrow As + Sn^{4+} + H_2O$ oxidising agent is –
(A) Sn^{4+} (B) Sn^{2+}
(C) As (D) $HAsO_4$

Q.23 Write the following ions in order of decreasing capacity to accept electrons.
$H^+, Mg^{2+}, K^+, Ag^+, Zn^{2+}$
(A) $Ag^+ > H^+ > Zn^{2+} > Mg^{2+} > K^+$
(B) $H^+ > Zn^{2+} > Mg^{2+} > K^+ > Ag^+$
(C) $K^+ > Mg^{2+} > Zn^{2+} > H^+ > Ag^+$
(D) $Mg^{2+} > Zn^{2+} > K^+ > Ag^+ > H^+$

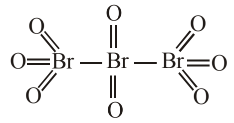
Q.24 Oxidation state of oxygen atom in potassium peroxide is
(A) zero (B) –1/2
(C) –1 (D) –2

Q.25 Identify the element which neither exhibits negative nor does the +ve oxidation state.
(A) Cs (B) Ne
(C) I (D) F

Q.26 In the reaction $3\,I_2 + 6\,NaOH \rightarrow NaIO_3 + 5\,NaI + 3H_2O$ oxidising agent is
(A) NaOH (B) $NaIO_3$
(C) I_2 (D) NaI

Q.27 Which of the following metal does not liberate H_2 from dilute H_2SO_4?
(A) Zn (B) Al
(C) Cu (D) Mg

Q.28 Oxidation number of bromine in sequence in Br_3O_8 is –



(A) +8, +6, +8 (B) +6, +4, +6
(C) 0, 0, 0 (D) +8, +4, +8

Q.29 The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following rules is not correct 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.

Q.30 Which of the following is a decreasing order of oxidation states of the central atoms?
(A) $PCl_5, HIO_4, Cr_2O_7^{2-}, Cl_2O$
(B) $Cr_2O_7^{2-}, Cl_2O, HIO_4, PCl_5$
(C) $HIO_4, Cr_2O_7^{2-}, PCl_5, Cl_2O$
(D) $Cr_2O_7^{2-}, HIO_4, Cl_2O, PCl_5$

EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)

NOTE : The answer to each question is a NUMERICAL VALUE.

Q.1 The oxidation number of Mn in the product of alkaline oxidative fusion of MnO_2 is

Q.2 The difference in the oxidation numbers of the two types of sulphur atoms in $Na_2S_4O_6$ is

Q.3 Consider a titration of potassium dichromate solution with acidified Mohr's salt solution using diphenylamine as indicator. The number of moles of Mohr's salt required per mole of dichromate is –

Q.4 An aqueous solution 6.3gm. of oxalic acid dihydrate is made upto 250 ml. The volume of 0.1 N NaOH required to completely neutralize 10 mol. of this solution is –

Q.5 6.90 gm of a metal carbonate were dissolved in 60 ml of 2(N) HCl. The excess acid was neutralized by 20 ml of 1(N) NaOH. The equivalent wt. of metal is

Q.6 The equiv. wt. of an element is 9. If it forms volatile chloride of vapour density 58.5. The approximate at wt. of the element is

Q.7 The equivalent wt. of element is 5.33. If it forms oxychloride which is isomorphous with CrO_2Cl_2. The atomic weight of element is-

Q.8 $(NH_4)_3PO_4$ get converted into NO_3^- and PH_3 on reacting with $KMnO_4$ in acidic medium. If 50ml of 0.2 M $(NH_4)_3PO_4$ solution reacts with 16ml of $KMnO_4$ solution, then normality of $KMnO_4$ solution is –

Q.9 The moles of ammonium sulphate needed to react with one mole of MnO_2 in acidic medium in a reaction giving $MnSO_4$ and $(NH_4)_2S_2O_8$ is

Q.10 The oxidation number of phosphorus in $MgNH_4PO_4$ is-

EXERCISE - 4 [PREVIOUS YEARS AIEEE / JEE MAIN QUESTIONS]

Q.1 MnO_4^- is good oxidising agent in different medium changing to –
$MnO_4^- \rightarrow Mn^{2+}$; $MnO_4^- \rightarrow MnO_4^{2-}$
$MnO_4^- \rightarrow MnO_2$; $MnO_4^- \rightarrow Mn_2O_3$
Changes in oxidation number respectively are –
(A) 1, 3, 4, 5 (B) 5, 4, 3, 2 **[AIEEE-2002]**
(C) 5, 1, 3, 4 (D) 2, 6, 4, 3

Q.2 Oxidation number of Cl in $CaOCl_2$ (bleaching powder) is
(A) Zero, since it contains Cl_2 **[AIEEE-2002]**
(B) –1, since it contains Cl^-
(C) +1, since it contains ClO^-
(D) +1 and –1 since it contains ClO^- and Cl^-

Q.3 Which of the following is a redox-reaction **[AIEEE-2002]**
(A) $2Na[Ag(CN)_2] + Zn \longrightarrow Na_2[Zn(CN)_4] + 2Ag$
(B) $BaO_2 + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2$
(C) $N_2O_5 + H_2O \longrightarrow 2HNO_3$
(D) $AgNO_3 + KI \longrightarrow AgI + KNO_3$

Q.4 In the coordination compound, $K_4[Ni(CN)_4]$, the oxidation state of nickel is – **[AIEEE-2003]**
(A) +1 (B) +2
(C) –1 (D) 0

Q.5 What would happen when a solution of potassium chromate is treated with an dilute nitric acid ?
(A) CrO_4^{2-} is reduced to +3 state of Cr **[AIEEE-2003]**
(B) CrO_4^{2-} is oxidized to +7 state of Cr
(C) Cr^{3+} and $Cr_2O_7^{2-}$ are formed
(D) $Cr_2O_7^{2-}$ and H_2O are formed

Q.6 The oxidation state of Cr in $[Cr(NH_3)_4Cl_2]^+$ is –
(A) +2 (B) +3 **[AIEEE-2005]**
(C) 0 (D) +1

Q.7 The oxidation state of chromium in the final product formed by the reaction between KI and acidified potassium dichromate solution is – **[AIEEE-2005]**
(A) +6 (B) +4
(C) +3 (D) +2

Q.8 Which of the following chemical reactions depicts the oxidizing behaviour of H_2SO_4 ? **[AIEEE 2006]**
(A) $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2H_2O$
(B) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$
(C) $2PCl_5 + H_2SO_4 \rightarrow 2POCl_3 + 2HCl + SO_2Cl_2$
(D) $2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$

Q.9 Consider the following reaction : **[JEE MAIN 2013]**

$$x\,MnO_4^- + yC_2O_4^{2-} + zH^+ \rightarrow xMn^{2+} + 2yCO_2 + \frac{z}{2}H_2O$$

The values of x, y and z in the reaction are, respectively
(A) 5, 2 and 16 (B) 2, 5 and 8
(C) 2, 5 and 16 (D) 5, 2 and 8

Q.10 In which of the following reactions H_2O_2 acts as a reducing agent? **[JEE MAIN 2014]**
(a) $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$
(b) $H_2O_2 - 2e^- \rightarrow O_2 + 2H^+$
(c) $H_2O_2 + 2e^- \rightarrow 2OH^-$
(d) $H_2O_2 + 2\,OH^- - 2e^- \rightarrow O_2 + 2H_2O$
(A) (a), (c) (B) (b), (d)
(C) (a), (b) (D) (c), (d)

Q.11 The equation which is balanced and represents the correct product(s) is: **[JEE MAIN 2014]**
(A) $[Mg(H_2O)_6]^{2+} + (EDTA)^{4-} \xrightarrow{\text{Excess NaOH}} [Mg(EDTA)]^{2+} + 6H_2O$
(B) $CuSO_4 + 4KCN \rightarrow K_2[Cu(CN)_4] + K_2SO_4$
(C) $Li_2O + 2KCl \rightarrow 2LiCl + K_2O$
(D) $[CoCl(NH_3)_5]^+ + 5H^+ \rightarrow Co^{2+} + 5NH_4^+ + Cl^-$

Q.12 In the reaction of oxalate with permaganate in acidic medium, the number of electrons involved in producing one molecule of CO_2 is : **[JEE MAIN 2019 (JAN)]**
(A) 10 (B) 2
(C) 1 (D) 5

Q.13 The highest possible oxidation states of uranium and plutonium, respectively, are : **[JEE MAIN 2019 (APRIL)]**
(A) 6 and 4 (B) 7 and 6
(C) 4 and 6 (D) 6 and 7

Q.14 Oxidation state of potassium in K_2O, K_2O_2 & KO_2 are respectively – **[JEE MAIN 2020 (JAN)]**
(A) +1, +1, +1 (B) +1, +2, +4
(C) +1, +2, +2 (D) +1, +4, +2

Q.15 Amongs the following which is redox reaction ? **[JEE MAIN 2020 (JAN)]**
(A) Combination of dinitrogen with dioxygen at 2000 K.
(B) Formation of ozone from atmosphereic oxygen in the presence of sunlight.
(C) Reaction of H_2SO_4 with NaOH
(D) Reaction of $[Co(H_2O)_6]Cl_3$ with $AgNO_3$.

Q.16 The compound that cannot act both as oxidising and reducing agent is : **[JEE MAIN 2020 (JAN)]**
(A) H_2O_2 (B) H_2SO_3
(C) HNO_2 (D) H_3PO_4

EXERCISE - 5 (PREVIOUS YEARS AIPMT/NEET EXAM QUESTIONS)

Choose one correct response for each question.

Q.1 The number of moles of $KMnO_4$ that will be needed to react with one mole of sulphite ion in acidic solution is
(A) 4/5 (B) 2/5 **[AIPMT 2007]**
(C) 1 (D) 3/5

Q.2 On the basis of the following E° values, the strongest oxidising agent is:
$[Fe(CN)_6]^{4-} \rightarrow [Fe(CN)_6]^{3-} + e^{-1}$; E° = – 0.35 V
$Fe^{2+} \rightarrow Fe^{3+} + e^{-1}$; E° = – 0.77 V **[AIPMT 2008]**
(A) $[Fe(CN)_6]^{3-}$ (B) $[Fe(CN)_6]^{4-}$
(C) Fe^{2+} (D) Fe^{3+}

Q.3 Number of moles of MnO_4^- required to oxidize one mole of ferrous oxalate completely in acidic medium will be **[AIPMT 2008]**
(A) 0.2 moles (B) 0.6 moles
(C) 0.4 moles (D) 7.5 moles

Q.4 Oxidation numbers of P in PO_4^{3-}, of S in SO_4^{2-} and that of Cr in $Cr_2O_7^{2-}$ are respectively: **[AIPMT 2009]**
(A) + 3, + 6 and + 5 (B) + 5, + 3 and + 6
(C) – 3, + 6 and + 6 (D) + 5, + 6 and + 6

Q.5 Which one of the elements with the following outer orbital configurations may exhibit the largest number of oxidation states? **[AIPMT 2009]**
(A) $3d^5 4s^1$ (B) $3d^5 4s^2$
(C) $3d^2 4s^2$ (D) $3d^3 4s^2$

Q.6 Among the following which is the strongest oxidising agent? **[AIPMT 2009]**
(A) Br_2 (B) I_2
(C) Cl_2 (D) F_2

Q.7 Oxidation states of P in $H_4P_2O_5$, $H_4P_2O_6$, $H_4P_2O_7$ are respectively : **[AIPMT (PRE) 2010]**
(A) +3, +5, +4 (B) +5, +3, +4
(C) +5, +4, +3 (D) +3, +4, +5

Q.8 Standard electrode potential of three metals X, Y and Z are –1.2 V, + 0.5 V and –3.0 V respectively. The reducing power of these metals will be **[AIPMT (PRE) 2011]**
(A) X > Y > Z (B) Y > Z > X
(C) Y > X > Z (D) Z > X > Y

Q.9 When Cl_2 gas reacts with hot and concentrated sodium hydroxide solution, the oxidation number of chlorine changes from : **[AIPMT (PRE) 2012]**
(A) Zero to +1 and zero to –5
(B) Zero to –1 and zero to +5
(C) Zero to –1 and zero to +3
(D) Zero to +1 and zero to –3

Q.10 In which of the following compounds, nitrogen exhibits highest oxidation state ? **[AIPMT (PRE) 2012]**
(A) N_2H_4 (B) NH_3
(C) N_3H (D) NH_2OH

Q.11 A mixture of potassium chlorate, oxalic acid and sulphuric acid is heated. During the reaction which element undergoes maximum change in the oxidation number ? **[AIPMT (PRE) 2012]**
(A) S (B) H
(C) Cl (D) C

Q.12 Standard reduction potentials of the half reactions are
$F_2(g) + 2e^- \rightarrow 2F^-(aq)$; E° = + 2.85 V
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$; E° = + 1.36 V
$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$; E° = + 1.09 V
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$; E° = + 0.53 V
The strongest oxidising and reducing agents respectively are **[AIPMT (MAINS) 2012]**
(A) F_2 and I^- (B) Br_2 and Cl^-
(C) Cl_2 and Br^- (D) Cl_2 and I_2

Q.13 The pair of compounds that can exist together is **[AIPMT 2014]**
(A) $FeCl_3$, $SnCl_2$ (B) $HgCl_2$, $SnCl_2$
(C) $FeCl_2$, $SnCl_2$ (D) $FeCl_3$, KI

Q.14 (a) $H_2O_2 + O_3 \rightarrow H_2O + 2O_2$
(b) $H_2O_2 + Ag_2O \rightarrow 2Ag + H_2O + O_2$
Role of hydrogen peroxide in the above reactions is respectively **[AIPMT 2014]**
(A) Oxidizing in (a) and reducing in (b)
(B) Reducing in (a) and oxidizing in (b)
(C) Reducing in (a) and (b)
(D) Oxidizing in (a) and (b)

Q.15 In acidic medium, H_2O_2 changes $Cr_2O_7^{-2}$ to CrO_5 which has two (– O – O –) bonds. Oxidation state of Cr in CrO_5 is **[AIPMT 2014]**
(A) +5 (B) +3
(C) +6 (D) –10

Q.16 Hot concentrated sulphuric acid is a moderately strong oxidizing agent. Which of the following reactions does not show oxidizing behaviour? **[NEET 2016 PHASE 1]**
(A) $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$
(B) $3S + 2H_2SO_4 \rightarrow 3SO_2 + 2H_2O$
(C) $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$
(D) $CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$

Q.17 It is because of inability of ns^2 electrons of the valence shell to participate in bonding that – **[NEET 2017]**
(A) Sn^{2+} is oxidising while Pb^{4+} is reducing.
(B) Sn^{2+} & Pb^{2+} are both oxidising & reducing.
(C) Sn^{4+} is reducing while Pb^{4+} is oxidising.
(D) Sn^{2+} is reducing while Pb^{4+} is oxidising

Q.18 The correct order of N-compounds in its decreasing order of oxidation states is **[NEET 2018]**
(A) HNO_3, NH_4Cl, NO, N_2 (B) HNO_3, NO, NH_4Cl, N_2
(C) HNO_3, NO, N_2, NH_4Cl (D) NH_4Cl, N_2, NO, HNO_3

Q.19 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 **[NEET 2018]**

	MnO_4^-	$C_2O_4^{2-}$	H^+
(A)	2	16	5
(B)	2	5	16
(C)	16	5	2
(D)	5	16	2

Q.20 Which of the following reactions are disproportionation reaction? **[NEET 2019]**

(a) $2Cu^+ \rightarrow Cu^{2+} + Cu^0$

(b) $3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$

(c) $2KMnO_4 \xrightarrow{\Delta} K_2MnO_4 + MnO_2 + O_2$

(d) $2MnO_4^- + 3Mn^{2+} + 2H_2O \rightarrow 5MnO_2 + 4H^+$

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

ANSWER KEY

EXERCISE - 1

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	A	B	D	D	C	D	B	D	A	D	A	D	A	D	D	B	C	D	C	D	B	C	A	C	D
Q	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50
A	A	C	C	B	B	B	A	D	D	D	B	D	D	C	A	A	D	C	C	D	D	A	A	A	A
Q	51	52	53	54	55	56	57	58	59	60	61	62	63	64	65	66	67	68							
A	D	B	C	C	C	A	B	B	B	B	A	B	B	C	D	D	B	A							

EXERCISE - 2

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
A	A	A	B	C	B	D	B	D	B	C	D	A	B	D	B	D	A	B	C	B	D	D	A	C	B	C	C	B	A	C

EXERCISE - 3

Q	1	2	3	4	5	6	7	8	9	10
A	6	5	6	40	39	27	32	10	2	5

EXERCISE - 4

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16
A	C	D	A	D	D	B	C	D	C	B	D	C	D	A	A	D

EXERCISE - 5

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
A	B	D	B	D	B	D	D	D	B	C	C	A	C	C	C	D	D	C	B	A

REDOX REACTIONS

TRY IT YOURSELF

(1) **(A).** N_A no. of electron will be removed by

$\frac{6.023\times10^{23}}{2.25\times10^{23}}\times16gm$ of metal M = 42.83 gm of metal M

$\therefore$ Equivalent wt. of metal is 42.83.

(2) **(B).** The equivalent wt. of $PO_4 = \frac{31\times4}{5\times4}=\frac{31}{5}$

$\therefore$ $62gm\ P_4 = \frac{62\times5}{31}$ equiv. of P_4 = 10 equiv. of P_4

The equiv. wt. of $HNO_3 = \frac{\text{Mol. wt}}{1}=\frac{63}{1}$

$\therefore$ The wt. of HNO_3 required = 10 × 63 = 630 gm.

(3) **(A).** In the above reaction $C_2O_4^{-2}$ acts as a reductant because it is oxidised to CO_2 as

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

$C_2O_4^{-2}$ reduces MnO_4^- to Mn^{+2} ion in solution.

(4) **(C).** Let the O.N. of Co be x

O.N. of NH_3 is zero.

O.N. of Cl is –1.

O.N. of Br is –1.

Hence, x + 6 (0) – 1 × 2 – 1 = 0

$\therefore$ x = +3

So, the oxidation number of cobalt in the given complex compound is +3.

(5) $N_2H_4\rightarrow(Y)+10e^-$ ($\because$ Y contains all N atoms)

$N_2^{2-}\rightarrow(2N)^x+10e^-$

–4 = 2x – 10

x = +3

$\therefore$ Oxidation state of N is Y is +3.

(6) **(C).** $Fe^{+2}\rightarrow Fe^{3+}+e^-$

$Cr_2^{6+}\rightarrow Cr^{+6}+6e^-$

$FeCr_2O_4\rightarrow Fe_2O_3\ K_2CrO_4+7e^-$

(7) **(B).** Redox titration

Eq. of $K_2C_2O_4.3H_2C_2O_4.4H_2O$ = Eq. of $KMnO_4$

$\frac{0.1\times V}{1000}=\frac{20\times0.05\times5}{1000}$; V = 50ml.

n factor of $K_2C_2O_4.3H_2C_2O_4.4H_2O$

For redox titration = 8

For acid base titration = 6

For acid base titration normality of

$K_2C_2O_4.3H_2C_2O_4.4H_2O=\frac{0.1}{8}\times6N$

Eq. of acid = Eq. of base

$\frac{0.1\times50}{8\times1000}\times6=\frac{1}{8}\times\frac{Vml}{1000}$; Vml = 30 ml.

(8) **(D).** Eq. of NaH_2PO_3 + Eq. of $NaHCO_3$ = Eq. of NaOH

$\frac{20\times0.1}{1000}\times1+\frac{40\times0.1}{1000}\times1=x$; $x=6\times10^{-3}$

(9) **(B).** O.N. of N in HNO_2 is +3

Max. O.N. of N is +5

Min. O.N. of N is –3

Thus, O.N. of N in HNO_2 can show an increase or decrease as the same may be. That is why HNO_2 acts as oxidant and reductant both.

O.N. of N in HNO_3 is +5, hence it can act only as an oxidant.

(10) **(C).** In the reaction, $P_2O_5\rightarrow H_4P_2O_7$

The O.N. of P in P_2O_5 is 2x + 5 (–2) = 0 or x = +5

The O.N. of P in $H_4P_2O_7$ is 4 (+1) + 2 (x) + 7 (–2) = 0

2x = 10 or x = +5

Since there is no change in O.N. of P, hence the above reaction is neither oxidation nor reduction.

(11) **(C).** O.N. of Fe in wustite is $\frac{290}{93}=2.15$

It is an intermediate value in between Fe(II) & Fe(III).

Let % of Fe (III) be a, then

2 × (100 – a) + 3a = 2.15 × 100

a = 15.05

$\therefore$ % of Fe (III) = 15.05%

(12) **(A).** $O_2+FeS_2\rightarrow FeO+SO_2$

Oxidation half reaction :

$S_2^{2-}\rightarrow 2S^{4+}+10e^-$ .... (1)

Reduction half reaction :

$4e^-+O_2\rightarrow 2O^{2-}$ .... (2)

So, on doing 2 × (1) + 5 (2)

$\therefore$ $2S_2^{2-}+5O_2\rightarrow 4S^{4+}+10O^{2-}$

$2FeS2+5O_2\rightarrow 4SO_2+2FeO$

$\because$ Since on molecule of FeS_2 liberates 10 electrons.

So, 2 moles of FeS_2 required to liberate 20 mole e^-.

CHAPTER-8 : REDOX REACTION

EXERCISE-1

(1) **(A).** Oxidation is a process in which hydrogen is removed or oxygen is added or loss of electron takes place or oxidation number increases.

(2) **(B).** At anode, oxidation takes place.

(3) **(D).** Reduction is addition of hydrogen and removal of oxygen.

(4) **(D).** When an element undergoes oxidation process, element converts into positive ion (electropositive element) by addition of electronegative element or by removal of electropositive element.

$$\underset{\text{(Element)}}{\underset{\downarrow}{X}} + \underset{\text{(Electronegative element)}}{\underset{\downarrow}{Y}} \rightarrow X^{+} + Y^{-} \rightarrow XY$$

(5) **(C).** $Cl_2 + OH^- \rightarrow Cl^- + ClO_4^- + H_2O$

$$\overset{0}{Cl_2} \rightarrow \overset{-1}{Cl^-} \text{ (Reduction)}$$

$$\overset{0}{Cl_2} \rightarrow \overset{+7}{ClO_4^-} \text{ (Oxidation)}$$

(6) **(D).** In reaction (A), the compound nitric oxide is formed by the combination of the elemental substances, nitrogen and oxygen; therefore, this is an example of combination redox reactions.

In reaction (B), hydrogen of water has been displaced by hydride ion into dihydrogen gas. Therefore, this may be called as displacement redox reaction.

The reaction (C) involves disproportionation of NO_2 (+4 state) into NO_2^- (+3 state) and NO_3^- (+5 state). Therefore reaction (C) is an example of disproportionation redox reaction.

(7) **(B).** $2Na + \overset{0}{S} \longrightarrow (\overset{+}{Na})_2\overset{-}{S}$

Gain of electrons (reduction)

(8) **(D).** Disproportionation reaction is a special type of redox reaction in which an element in one oxidation state is simultaneously oxidised and reduced.

Oxidation

$$2\overset{+4}{NO_2} + 2OH^- \rightarrow \overset{+3}{NO_2^-} + \overset{+5}{NO_3^-} + H_2O$$

Reduction

(9) **(A).** $CuO + \underbrace{2\overset{-3}{NH_3} \rightarrow \overset{0}{N_2}}_{-6e^-} + 3H_2O + 3Cu$

(10) **(D).** $2Na + Cl_2 \longrightarrow 2\overset{+}{Na}\overset{-}{Cl}$

Gain of e^- = 2

Loss of electron = 2

(11) **(A).** $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

In this reaction Sn^{2+} change in Sn^{++} it is called an oxidation reaction.

(12) **(D).** $MnO_4^- \rightarrow Mn^{2+}$

In this reaction $5e^-$ are needed for the reduction of Mn^{2+} as: $MnO_4^- + 5e^- \rightarrow Mn^{2+}$.

(13) **(A).** When a strip of metallic zinc is placed in an aqueous solution of $Cu(NO_3)_2$, zinc appears as ions (Zn^{2+}).

$$Zn(s) + \overset{2+}{Cu}(NO_3)_2(aq) \longrightarrow \overset{2+}{Zn}(NO_3)_2 + Cu(s)$$

$$Zn \xrightarrow{-2e^-} Zn^{2+}$$

(14) **(D).** $\overset{0}{Zn} + 2\overset{+1}{Ag}CN \rightarrow 2\overset{0}{Ag} + \overset{+2}{Zn}(CN)_2$

$$\overset{0}{Zn} \rightarrow \overset{+2}{Zn}(CN)_2 \text{ (Oxidation)}$$

$$\overset{+1}{Ag}CN \rightarrow \overset{0}{Ag} \text{ (Reduction)}$$

(15) **(D).** $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

is not a redox reaction as it does not involve any change in oxidation number. It is an example of double decomposition reaction.

(16) **(B).** $4\overset{0}{Na} + O_2 \longrightarrow 2\overset{+}{Na}_2O$

Loss of electron (oxidation)

In this reaction, Na converts into ion (Na^+) and Na donates electrons to oxygen atoms, So, Na behaves as reducing agent.

(17) **(C).** $2\overset{0}{Na} + 2\overset{+1}{H_2}O \rightarrow 2Na\overset{+1}{O}H + \overset{0}{H_2}$

Reduced

Oxidised

(18) **(D).** Let O. No. of C in $C_{12}H_{22}O_{11}$ be x

$\therefore \quad 12 \times x + 22 + 11(-2) = 0, x = 0$

(19) **(C).** A reagent which lowers the oxidation number of an element in a given substance is reducing agent or reductant.

(20) **(D).** O.N of O in $KO_3 : 1 + 3x = 0$

$\therefore \quad x = -1/3 = -0.33$

O.N. of O in $Na_2O_2 : 2 \times 1 + 2x = 0 \quad \therefore \quad x = -1$

(21) **(B).** In $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Oxidation number of Zn changes from zero to +2. Thus Zn is oxidised from Zn to Zn^{2+}

(22) **(C).** $\overset{+1}{Ag_2}O + H_2O + 2e^- \rightarrow 2\overset{0}{Ag} + 2OH^-$

In the given reaction silver is reduced by a change in oxidation number from +1 (Ag_2O) to zero (Ag).

(23) **(A).** $\overset{+1}{K}\overset{+1}{H_2}\overset{x}{P}\overset{-2}{O_2}$

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

(24) **(C).** Let the oxidation no. of S is 'a'

O.N. of $CH_3 = +1$

O.N. of $O = -2$

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

Hence the oxidation no. of S in Dimethyl sulphoxide is zero.

(25) **(D).** Na N_3
$1 + 3x = 0 \Rightarrow x = -1/3$
So, oxidation number of nitrogen in NaN_3 is $-1/3$.

(26) **(A).** In OF_2
$x + 2(-1) = 0, \; x = +2$
Oxidation number of oxygen in $OF_2 = +2$.

(27) **(C).** $2 \times$ No. of e^- losses = Oxidation number
$2 \times 3e^- = +6$.

(28) **(C).** In $H_2S_2O_8$, two O atoms form peroxide linkage.

$$\begin{array}{c} \quad\quad O \quad\quad\quad\quad\quad O \\ \quad\quad \uparrow \quad\quad\quad\quad\quad \uparrow \\ H-O-S-O-O-S-O-H \\ \quad\quad \downarrow \quad\quad\quad\quad\quad \downarrow \\ \quad\quad O \quad\quad\quad\quad\quad O \end{array}$$

$2 \times 1 + 2a + 6(-2) + 2(-1) = 0 \qquad \therefore\ a = +6$
Thus the O.N. of S in $H_2S_2O_8$ is $+6$

(29) **(B).** Oxygen shows + 2 oxidation state in F_2O.
As F most electronegative element, it always has an oxidation number $= -1$

(30) **(B).** $\overset{*}{N}H_2OH$
$x + 2(+1) - 2 + 1 = 0; \; x + 2 - 2 + 1 = 0; \; x = -1$.

(31) **(B).** $\overset{*}{C}H_3-Cl$
$x + 3(+1) + (-1) \times 1 = 0 \Rightarrow x = -2$.

(32) **(A).** $\overset{*}{S}_2Cl_2$
$2x + 2(-1) = 0; \; 2x - 2 = 0 \Rightarrow x = +1$

(33) **(D).** $\overset{*}{Fe}_3O_4$
$3x + (-8) = 0; \; 3x - 8 = 0 \Rightarrow x = 8/3$.

(34) **(D).** $K_4[\overset{*}{Ni}(CN)_4]$
$4 \times (+1) + x + 4 \times (-1) = 0 \Rightarrow x = 0$

(35) **(D).** Fluorine is the most electronegative element in the periodic table so it never shows positive oxidation state.

(36) **(B).** Let the oxidation number of I in $IPO_4 = x$
Oxidation number of $PO_4 = -3$
$x + (-3) = 0 \Rightarrow x = +3$

(37) **(D).** $\overset{x}{N_3}\overset{+1}{H}$ i.e., $3(x) + 1(+1) = 0 \Rightarrow x = -1/3$

(38) **(D).** F exhibits only negative oxidation state of -1.

(39) **(C).** Sum of the oxidation numbers of atoms must be zero.

(40) **(A).** $\overset{+2}{Ca}\overset{+4}{C}\overset{-2}{O_3} \rightarrow \overset{+2}{Ca}\overset{-2}{O} + \overset{+4}{C}\overset{-2}{O_2}$
No change in oxidation number, hence it can't be a redox reaction.

(41) **(A).** $\overset{+3}{P}\overset{-1}{Cl_3} + \overset{0}{Cl_2} \rightarrow \overset{+5}{P}\overset{-1}{Cl_5}$
Oxidation state of P is changing from (+3) to (+5) i.e., PCl_3 is oxidised and hence acts as a reductant.

(42) **(D).** $\overset{x}{N}\overset{-2}{O_3^-}$ i.e., $1(x) + 3(-2) = -1 \Rightarrow x = +6 - 1 = +5$.

(43) **(C).** With reduction, participation of atom electron decreases, therefore, Oxidation number also decreases.

(44) **(C).** $\overset{x}{C}\overset{+1}{H_2}\overset{-2}{O}$ i.e., $x + 2(+1) + 1(-2) = 0$
$\Rightarrow \; x = 0$

(45) **(D).** $x - 2 = 0 \Rightarrow x = +2$.

(46) **(D).** In elemental state, oxidation state is zero.

(47) **(A).** $[\overset{x}{Cl}\overset{-2}{O_3}]^- : x + 3(-2) = -1 \Rightarrow x = +5$

(48) **(A).** In simple ions, oxidation number is equal to the charge.

(49) **(A).** $\overset{+2}{Ba_2}\overset{x}{Xe}\overset{-2}{O_6} : 2(+2) + x + 6(-2) = 0 \Rightarrow x = +8$

(50) **(A).** In metal carbonyl, metal is in zero oxidation state.

(51) **(D).** Balanced equation is
$2K_2CrO_4 + 2HCl \rightarrow K_2Cr_2O_7 + 2KCl + H_2O$

(52) **(B).** The balanced equation is
$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
Ratio of the coefficients of CO_2 and H_2O is 4 : 6 or 2 : 3.

(53) **(C).**
$\overset{0}{I_2} \rightarrow \overset{+5}{IO_3}$ ..... (i) Oxidation
$\overset{0}{I} \rightarrow \overset{-1}{I^-}$ ..... (ii) Reduction
Balancing atoms of Iodine on two sides,
we get, $I_2 \rightarrow 2IO_3^-$
or, $I_2 + 12OH \rightarrow 2IO_3^- + 6H_2O$
Balancing change,
$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-$
$I_2 \rightarrow 2I^-$
and $(I_2 + 2e^- \rightarrow 2I^-) \times 5$

Adding,
$6I_2 + 12OH^- \rightarrow 2IO_3 + 10I^- + 6H_2O$
or, $3I_2 + 6OH^- \rightarrow IO_3^- + 5I^- + 3H_2O$
It is balanced equation.

(54) **(C).** The values of x, y, z are 8, 4, 4 respectively hence the reaction is
$H_2SO_4 + 8HI \rightarrow H_2S + 4I_2 + 4H_2O$

(55) **(C).** The correctly balanced half reaction is –
$Cr_2O_7^{-2} + 14H^+ \rightarrow 2Cr^{+3} + 7H_2O - 6e^-$
It is a reduction half reaction in balancing the equation by ion–electron method.

(56) **(A).** Out of given choices x = 2, y = 5 and z = 6 makes the equation perfectly balanced.
$2MnO_4^- + 5H_2SO_4 \rightarrow 2Mn^{2+} + 5H_2O + 9O_2 + 6e^-$

(57) **(B).** $5H_2O_2 + 2ClO_2 + 2OH^- \rightarrow 2Cl^- + 5O_2 + 6H_2O$

(58) **(B).** $4Na + O_2 \rightarrow 2Na_2O$

(59) **(B).** Out of given option only x = 2 and y = 6 makes the equation balanced.
$2Cu + 6HNO_3 \rightarrow 2Cu(NO_3)_2 + NO + NO_2 + 3H_2O$.

(60) **(B).** The oxidation state of Bi in BiO_3^- is +5 and it reduces to Bi^{3+}. $\therefore$ The value of x is 2.

(61) **(A).** F_2 is the strongest oxidising agent with highest reduction potential.

(62) **(B).** A negative E^{Θ} means that the redox couple is a stronger reducing agent than H^+/H_2 couple.
$-E^{\Theta}$ = strong reducing agent.
$+E^{\Theta}$ = weak reducing agent.

(63) **(B).** An electronegative element in its lowest oxidation state acts as powerful reducing agent.

(64) **(C).** Since Ag has higher reduction potential than Cu, Ag will not reduce Cu^{2+} to Cu. Cu can reduce Ag^+ to Ag.

(65) **(D).** $\overset{+3}{Al_2O_3}$ is reduced to $\overset{0}{Al}$.

(66) **(D).** More positive the value of E°, greater is the tendency of the species to get reduced. On the basis of the given E° values, Br_2 is having highest E° value (+1.90) hence, Cu will easily reduce Br_2.

(67) **(B).** Lower the value of reaction potential, stronger is the reducing agent.

(68) **(A).** Out of Cl, Br and I, I has lower ionisation energy and lowest reduction potential.

EXERCISE-2

(1) **(A).** When $Cr_2O_7^{-2}$ is converted into CrO_4^{-2} the change in oxidation number of Cr is zero.
$$\underset{+6}{Cr_2O_7^{-2}} \rightarrow \underset{+6}{CrO_4^{-2}}$$

(2) **(A).**

(A) $\overset{+2}{CO} \rightarrow \overset{+4}{CO_2}$ (B) $\overset{+2}{CuO} \rightarrow \overset{+2}{CuCl_2}$

(C) $\overset{+1}{H_2O} \rightarrow \overset{0}{H_2}$ (D) $\overset{0}{C} \rightarrow \overset{+4}{CO_2}$

Only in (C), O.N. of hydrogen decreases from + 1 to 0 and hence H_2O gets reduced to H_2.

(3) **(B).** Oxidation number of N in HNO_2 is + 3
Maximum O.N. of N is + 5
Minimum O.N. of N is – 3
Thus O.N. of N in HNO_2 can show an increase or decrease as the case may be. That is why HNO_2 acts as oxidant and reductant both.
Oxidation number of N in HNO_3 is + 5,
Hence it can act only as an oxidant.

(4) **(C).** In the reaction $P_2O_5 \rightarrow H_4P_2O_7$
The Oxidation number of P in P_2O_5 is
$2x + 5(-2) = 0$ or $x = +5$
The oxidation number of P in $H_4P_2O_7$ is
$4(+1) + 2(x) + 7(-2) = 0;\ 2x = 10;\ x = +5$
Since there is no change in oxidation number of P, hence the above reaction is neither oxidation nor reduction.

(5) **(B).** $YBa_2\overset{*}{Cu_3}O_7$
$3 + 2 \times 2 + 3x - (2 \times 7) = 0$
$3 + 4 + 3x - 14 = 0$
$3x = 7 \Rightarrow x = 7/3$

(6) **(D).** Since a metal with lower electrode potential is a stronger reducing agent, Mg can displace all the given metals, Al can displace all metals except Mg.
Zn can displace all metals except Mg and Al. Fe can displace only Cu. The order in which they can displace each other from their salt solutions is Mg, Al, Zn, Fe, Cu.

(7) **(B).** $2Al^o \rightarrow Al_2^{+3} + 6e^-$ .... (1)
$8e + Fe_3^{+8/3} \rightarrow 3Fe^o$ .... (2)
Multiplying Eq.(1) by 4 & Eq. (2) by 3, then on addition
$8Al^o \rightarrow 4Al_2^{+3} + 24e$
$24e + 3Fe_3^{+8/3} \rightarrow 9Fe^o$

$8Al^o + 3Fe_3^{+8/3} \rightarrow 9Fe^o + 4Al_2^{+3}$

or $8Al + 3Fe_3O_4 \rightarrow 4Al_2O_3 + 9Fe$
Therefore, it is clear that total number of electrons transferred during change = 24

(8) **(D).** The balanced redox reaction given above can be written as : $10FeC_2O_4 + 6KMnO_4 + 24H_2SO_4 \rightarrow 5Fe_2(SO_4)_3 + 20CO_2 + 6MnSO_4 + 3K_2SO_4 + 24H_2O$
So the value of x = 6 and y = 6

(9) **(B).** The reaction are
$MnO^-_4 + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$
$A^{+n} + 3H_2O \rightarrow AO_3^- + 6H^+ + (5-n)e^-$
Amount of electrons involved in the given amount of $MnO_4^- = 5 \times 1.61 \times 10^{-3}$ mol.
Equating these two we get
$5 \times 1.61 \times 10^{-3} = (5 - n)\,2.68 \times 10^{-3}$
$\therefore$ n = 2 (approx.)

(10) **(C).** $\overset{*}{Cr}O_4^{2-}$ $\qquad$ $\overset{*}{Cr_2}O_7^{2-}$
$x + [(-2) \times 4] = -2$ $\qquad$ $2x + (-2) \times 7 = -2$
$x = 8 - 2 = +6$ $\qquad$ $2x = 14 - 2 = 12$,
$x = \frac{12}{2} = +6$
In this reaction oxidation and reduction are not involved because there is no change in oxidation number.

(11) **(D).** HClO is the strongest oxidising agent.
The correct order of oxidising power is
$\overset{+1}{HClO} > \overset{+3}{HClO_2} > \overset{+5}{HClO_3} > \overset{+7}{HClO_4}$.

(12) **(A).** When sulphur dioxide is react with H_2S here SO_2 act as an oxidising agent and H_2S act as reducing agent.

(13) **(B).** $\overset{+1}{H_2}\overset{x}{O_2}$ i.e., $2(+1) + 2(x) = 0 \Rightarrow x = -1$

(14) **(D).** $H_2O + \underset{0}{Br_2} \longrightarrow \underset{+1}{HOBr} + \underset{-1}{HBr}$
In the above reaction the oxidation number of Br_2 increases from zero (in Br_2) to +1 (in HOBr) and decrease from zero (Br_2) to – 1 (in HBr). Thus Br_2 is oxidised as well as reduced.

(15) **(B).** $[\overset{*}{Cr}(H_2O)_4Cl_2]^+$; $x + 0 + 2(-1) = +1$; $x - 2 = +1$
$x = +3$ for Cr in complex.

(16) **(D).** $(NH_4)_2SO_4 \rightleftharpoons 2NH_4^+ + SO_4^{--}$

$\overset{*}{N}H_4^+$: $x + 4 = +1$; $x = 1 - 4 = -3$.

(17) **(A).** Mn shows + 7 oxidation state in MnO_4^{-1}

$\overset{*}{Mn}O_4^{-1}$: $x + (-2 \times 4) = -1 \Rightarrow x - 8 = -1 \Rightarrow x = -1 + 8 = +7$.

(18) **(B).** In complex $[Pt(C_2H_4)Cl_3]^-$ Pt have + 2 oxidation state. $x + (-3) = -1$

(19) **(C).** Higher is the reduction potential stronger is the oxidising agent. Hence in the given options. MnO_4^- is strongest oxidising agent.

(20) **(B).** $14H^+ + Cr_2O_7^{2-} + 3Ni \rightarrow 2Cr^{3+} + 7H_2O + 3Ni^{2+}$
i.e., Ni acts as a reducing agent.

(21) **(D).** The balanced reaction is
$2MnO_4^- + 5H_2O_2 + 6H^+ \rightarrow 2Mn^{2+} + 5O_2 + 8H_2O$

(22) **(D).** Oxidizing agent, itself, undergoes reduction during a redox reaction

$$\underbrace{\overset{+7}{HAsO_4} + Sn^{2+} \longrightarrow \overset{0}{As}}_{\text{Reduction}} + Sn^{4+} + H_2O$$

Here $HAsO_4$ is acting as oxidizing agent.

(23) **(A).** Based on the values of standard reduction potentials the sequence is followed.

(24) **(C).** $\overset{+1}{K_2}\overset{x}{O_2}$ i.e., $2(+1) + 2(x) = 0 \Rightarrow x = -1$.

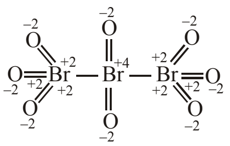
Alternatively, oxidation number of oxygen in peroxides is –1.

(25) **(B).** The oxidation state of Ne is zero. It exhibits neither negative nor positive oxidation states.

(26) **(C).** $3\overset{0}{I_2} + 6NaOH \rightarrow Na\overset{+5}{I}O_3 + 5Na\overset{-1}{I} + 3H_2O$

In the given reaction, I_2 is reduced thus acts as oxidising agent.

(27) **(C).** Cu is below hydrogen in the electrochemical series, so cannot liberate H_2 from dilute H_2SO_4,
$E°(Cu^{2+}/Cu) = +0.34$

(28) **(B).** 

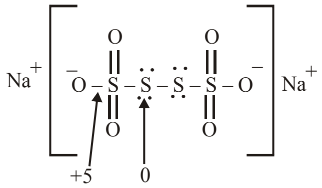
Oxidation states of three bromines are + 6, + 4, + 6

(29) **(A).** Oxidation number of hydrogen may be +1 or –1 or 0. For example, in HX it is +1, in hydrides (MH), it is –1 and in H_2 it is 0.

(30) **(C).** $Cl_2O : 2x + 1(-2) = 0 \Rightarrow x = +1$
$Cr_2O_7^{2-} : 2x + 7(-2) = -2 \Rightarrow x = +6$
$HIO_4 : +1 + x + 4(-2) = 0 \Rightarrow x = +7$
$PCl_5 : x + (-5) = 0 \Rightarrow x = +5$

EXERCISE-3

(1) **6.** $2MnO_2 + 4KOH + O_2 \xrightarrow{\text{fusion}} 2K_2MnO_4 + 2H_2O$
Let the oxidation state of Mn in MnO_4^{2-} is x.
So, $x + 4(-2) = -2$ or $x = 6$

(2) **5.** 

(3) **6.** $Cr_2O_7^{2-} + Fe^{2+} \longrightarrow Fe^{3+} + Cr^{3+}$
n factor of $Cr_2O_7^{2-} = 6$
n factor of $Fe^{2+} = 1$
So to reduce one mole of dichromate 6 moles of Fe^{2+} are required.

(4) **40.** Eq. wt. of $H_2C_2O_4.2H_2O$ = Eq. wt. of NaOH
Strength of $H_2C_2O_4.2H_2O$ (in g/L)

$$= \frac{6.3}{250/1000} = 25.2\,g/L$$

Normality of $H_2C_2O_4.2H_2O$

$$= \frac{\text{Strength}}{\text{Eq. wt}} = \frac{25.3}{126/2} = 0.4\,N$$

$$\underset{(H_2C_2O_4.2H_2O)}{N_1V_1} = \underset{(NaOH)}{N_2V_2}$$
$$0.4 \times 10 = 0.1 \times V_2$$

$$V_2 = \frac{0.4 \times 10}{0.1} = 40ml.$$

(5) **39.** Equiv. of HCl taken $= 60 \times 2 \times 10^{-3}$
Equiv. of HCl present after the reaction $= 20 \times 1 \times 10^{-3}$
$\therefore$ Equiv. of HCl utilized $= 100 \times 10^{-3}$
$\therefore$ 100×10^{-3} equiv. of metal carbonate = 6.90 gm.

$\therefore$ 1 equiv. of metal carbonate $= \dfrac{6.90}{10^{-1}} = 69$ gm.

$\therefore$ equiv. wt. of metal = 69 – 30 = 39
[because equiv. wt. of carbonate = 30]

(6) **27.** Let, the molecular formula of the chloride is MCl_x and at. wt. of the element is a
$\therefore$ $ax + x + 35.5 = 58.5 \times 2$

$$x = \frac{58.5 \times 2}{44.5} = 2.63$$

The nearest whole no. of 2.63 = 3
$\therefore$ Approximate at. wt. of the element $= 9 \times 3 = 27$

(7) **32.** The O.N. of Cr in CrO_2Cl_2 is +6.
$\therefore$ The O.N. of the element is also 6.
$\therefore$ The at. wt. of the element $= 5.33 \times 6 = 32$

(8) **10.** $(NH_4)_3PO_4 \xrightarrow[16ml]{KMnO_4} NO_3^- + PH_3$
–3 +5 +5 –3
+24 – 8 = 16
$50 \times 0.2M$

n-factor of $KMnO_4$ in acidic medium = 5
n-factor of $(NH_4)_3PO_4 = 16$
Eq. of $(NH_4)_3PO_4$ = Eq. of $KMnO_4$

$$\frac{0.2 \times 50}{1000} \times 16 = \frac{N \times 16}{1000} \Rightarrow N = 10$$

(9) **2.** $\overset{+4}{Mn}O_2 \longrightarrow \overset{+2}{Mn}SO_4 \quad (n=2)$

and $2SO_4^{2-} \longrightarrow S_2O_8^{2-} + 2e^- \quad (n=1)$

Let the moles of ammonium sulphate reacting with 1 mole of MnO_2 be x.
As, Equivalents of MnO_2 = Equivalents of $(NH_4)_2SO_4$
$\therefore$ $1 \times 2 = x \times 1, x = 2]$

(10) **5.** $MgNH_4PO_4$. The net charge is 0.
PO_4^{-3} charge is because Mg^{+2}, NH_4^{+1} neutralize.
So, P has the oxidation number of +5.

EXERCISE-4

(1) **(C).** $MnO_4^- \longrightarrow Mn^{+2} \xrightarrow{\text{change in 0.5}} 5$
$\longrightarrow \overset{+6}{Mn}O_4^{2-} \longrightarrow 1$
$\longrightarrow \overset{+4}{Mn}O_2 \longrightarrow 3$
$\longrightarrow \overset{+3}{Mn_2}O_3 \longrightarrow 4$

(2) **(D).** $CaOCl_2$ contains Ca^{+2}, OCl^- and Cl^-
Cl has O.N. = – 1, in Cl^-, whiles its O.N. in OCl^- part is +1.

(3) **(A).** $2NaAg(CN)_2 + Zn \longrightarrow Na_2Zn(CN)_4 + 2Ag$

(4) **(D).** $K_4[Ni(CN)_4]$
$4 \times (+1) + x + 4(-1) = 0$
$4 + x - 4 = 0 \ ; \ x = 0$

(5) **(D).** $2CrO_4^{2-} + 2H^+ \longrightarrow Cr_2O_7^{2-} + H_2O$

(6) **(B).** $[Cr(NH_3)_4Cl_2]^+$
$x + 4 \times 0 + 2 \times (-1) = +1 \ ; x - 2 = +1 \ ; \ x = +3$

(7) **(C).** $Cr_2O_7^{2-} + 14H^+ + 6I^- \longrightarrow 2Cr^{+3} + 7H_2O + 3I_2$

(8) **(D).** We know that oxidizing agent is the one whose O.N. decreases during the reaction in case of H_2SO_4 (O.N. of S = + 6), if charges to SO_2 (O.N. of S = + 4).

$$2HI + H_2\overset{+6}{S}O_4 \longrightarrow I_2 + \overset{+4}{S}O_2 + 2H_2O$$

(9) **(C).** $MnO_4^- + C_2O_4^{2-} + H^+ \rightarrow Mn^{2+} + CO_2 + H_2O$
$v_f = 1\,(7-2)$ $v_f = 2(3-2)$
$= 5$ $= 2$
$\therefore$ Balanced Equation :
$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$
So, x = 2, y = 5 & z = 16.

(10) **(B).** A reducing agent loses electrons during redox reaction. Hence (b, d) is correct.

(11) **(D).** Equation – 1 is not balanced w.r.t. charge.
Equation – 2 gives $K_3[Cu(CN)_4]$ as product.
Equation – 3 reaction is unfavourable in the forward direction (K_2O is unstable, while Li_2O is stable).
Equation – 4 is correct & balanced.

(12) **(C).** $2\overset{+7}{Mn}O_4 + 5C_2O_4^{2-} + 16H^+ \rightarrow 2\overset{+2}{Mn^{2+}} + 10CO_2 + 8H_2O$
10 e^- trans for 10 molecules of CO_2 so per molecule of CO_2 transfer of e^- is '1'.

(13) **(D).** The highest oxidation state of U and Pu is 6+ and 7+ respectively.

(14) **(A).** Potasisum has an oxidation of +1 (only) in combined state.

(15) **(A).** $N_2 + O_2 \rightarrow 2NO$
$3O_2 \rightarrow 2O_3$
$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
$AgNO_3 + NaCl \rightarrow NaNO_3 + AgCl$

(16) **(D).**
(i) H_2O_2 act as oxidising agent as well as reducing agent depending on condition.
(ii) H_2SO_3 act as oxidising agent as well as reducing agent depending on condition.
(iii) HNO_2 act as oxidising agent as well as reducing agent depending on condition.
(iv) H_3PO_4 can not act both as oxidising and reducing agent.
H_3PO_4 can act as only oxidising agent.
$H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{3-}$

EXERCISE-5

(1) **(B).** The balance chemical equation is

$$2MnO_4^- + 6H^+ + 5SO_3^{--} \longrightarrow 2Mn^{++} + 5SO_4^{--} + 3H_2O$$

From the equation it is clear that
Moles of MnO_4^- require to oxidise 5 moles of SO_3^{--} are 2.
Moles of MnO_4^- require to oxidise 1 moles of SO_3^{--} are 2/5.

(2) **(D).** Fe^{+3} has highest SRP & thus behaves as strongest oxidising agent.

(3) **(B).** $3MnO_4^- + 5C_2O_4^{2-} + 24H^+$
$\rightarrow 3Mn^{2+} + 10CO_2 + 12H_2O$
Number of moles of MnO_4^- required to oxidise 5 moles of $C_2O_4^{2-} = 3$
Number of moles of MnO_4^- required to oxidise 1 mole of $C_2O_4^{2-} = 3/5 = 0.6$ mol

(4) **(D).** Let the oxidation state of P in PO_4^{3-} is x.
$x + 4(-2) = -3 \ ; x - 8 = -3 \ ; x = +5$
Let the oxidation state of S in SO_4^{2-} is y.
$y + 4(-2) = -2 \ ; y - 8 = -2 \ ; y = +6$
Let the oxidation state of Cr in $Cr_2O_7^{2-}$ is z.
$2z + 7(-2) = -2 \ ; 2z - 14 = -2 \ ; z = +6$

(5) **(B).** $3d^3 4s^2 \Rightarrow$ O.S. = 3 + 2 = 5
$3d^5 4s^1 \Rightarrow$ O.S. = 5 + 1 = 6
$3d^5 4s^2 \Rightarrow$ O.S. = 5 + 2 = 7
$3d^2 4s^2 \Rightarrow$ O.S. = 2 + 2 = 4

(6) **(D).** Fluorine is the most electronegative element because electronegativity decreases on moving down the group. Hence, it gets reduced readily into F^- ion and is the strongest oxidising agent.

(7) **(D).** Oxidation state of H is + 1 and that of O is –2.
Let the oxidation state of P in the given compounds is x. In $H_4P_2O_5$, $(+1)\times 4 + 2\times x + (-2)\times 5 = 0$
$4 + 2x - 10 = 0 ; 2x = 6 \therefore x = +3$
In $H_4P_2O_6$, $(+1)\times 4 + 2\times x + (-2)\times 6 = 0$
$4 + 2x - 12 = 0 \quad \therefore x = +4$
In $H_4P_2O_7$, $(+1)\times 4 + 2\times x + (-2)\times 7 = 0$
$4 + 2x - 14 = 0 \quad \therefore x = +5$

(8) **(D).** Z > X > Y; higher the reduction potential lesser the reducing power.

(9) **(B).** $\overset{(0)}{Cl_2} + OH^- \rightarrow Na\overset{(-1)}{Cl} + Na\overset{(+5)}{Cl}O_3$
This is an example of disproportionation reaction and oxidation state of chlorine changes from 0 to –1& +5.

(10) **(C). Compound Oxidation number of nitrogen**
N_2H_4 $\quad 2x + 4(+1) = 0 \Rightarrow 2x + 4 = 0 \Rightarrow x = -2$
NH_3 $\quad x + 3(+1) = 0 \Rightarrow x = -3$
N_3H $\quad 3x + 1(+1) = 0 \Rightarrow 3x + 1 = 0 \Rightarrow x = -1/3$
NH_2OH $\quad x + 2 + 1(-2) + 1 = 0$
$\Rightarrow x + 1 = 0 \Rightarrow x = -1$

(11) **(C).** $\overset{+1\,+5\,-2}{KClO_3} + \overset{+1\,+3\,-2}{H_2C_2O_4} + \overset{+1\,+6\,-2}{H_2SO_4}$
$\rightarrow \overset{+1\,+6\,-2}{K_2SO_4} + \overset{+1\,-1}{KCl} + \overset{+4\,-2}{CO_2} + \overset{+1\,-2}{H_2O}$
Maximum change in oxidation number is observed in Cl (+5 to –1).

(12) **(A).** Higher the value of standard reduction potential, stronger will be the oxidising agent. Therefore, F_2 will act as strongest oxidising agent.
Similarly, lower the value of standard reduction potential, stronger will be the reducing agent. Therefore, I^- will act as strongest reducing agent.

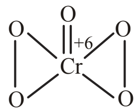
(13) **(C).** $\underset{OA}{FeCl_3} + \underset{RA}{SnCl_2} \rightarrow SnCl_4 + FeCl_2$
$\underset{OA}{HgCl_2} + \underset{RA}{SnCl_2} \rightarrow SnCl_4 + Hg_2Cl_2$
$\underset{OA}{FeCl_3} + \underset{RA}{KI} \rightarrow I_2 + FeCl_2$
$FeCl_2$ and $SnCl_2$ pair can exist together because $FeCl_2$ and $SnCl_2$ both are act as reducing agent.

(14) **(C).** Hydrogen peroxide generally act as an oxidising agent but in the presence of strong oxidising agent like $KMnO_4$, $Kr_2Cr_2O_7$, Halogen's and ozone, tollens reagent, it act as a reducing agent.

(a) $\underset{\text{R.A.}}{H_2O_2} + \underset{\text{O.A.}}{O_3} \longrightarrow H_2O + 2O_2$
$H_2O_2 \rightarrow H_2O\ [O]$; $O_3 \rightarrow [O]\ O_2$

(b) $\underset{\text{R.A.}}{H_2O_2} + \underset{\text{O.A.}}{Ag_2O} \longrightarrow H_2O + 2Ag + O_2$
$H_2O_2 \rightarrow H_2O\ [O]$; $Ag_2O \rightarrow 2Ag\ [O]$

(15) **(C).** CrO_5 has butterfly structure having two peroxo bonds.



Peroxo oxygen has –1 oxidation state.
Let oxidation state of Cr be x.
$CrO_5 : x + 4(-1) + 1(-2) = 0 \Rightarrow x = +6$

(16) **(D).** $CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$
In this reaction there is no change in oxidation state of any atom.

(17) **(D).** $Sn^{+2} \rightarrow Sn^{+4}$; (R.A) $Sn^{+2} < Sn^{+4}$ Stability order
$Pb^{+4} \rightarrow Pb^{+2}$; (O.A) $Pb^{+2} > Pb^{+4}$ Stability order
(Inert pair effect)

(18) **(C).** $\overset{+5}{HNO_3}, \overset{+2}{NO}, \overset{0}{N_2}, \overset{-3}{NH_4Cl}$.
The correct option is (C).

(19) **(B).** $\overset{+7}{MnO_4^-} + \overset{+3}{C_2O_4^{2-}} + H^+ \longrightarrow Mn^{2+} + \overset{+4}{CO_2} + H_2O$
Reduction ($MnO_4^- \rightarrow Mn^{2+}$); Oxidation ($C_2O_4^{2-} \rightarrow CO_2$)

n-factor of $MnO_4^- = 5$
n-factor of $C_2O_4^{2-} = 4$
Ratio of n-factors of MnO_4^- and $C_2O_4^{2-}$ is 5 : 2
So, molar ratio in balanced reaction is 2 : 5
The balanced equation is
$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

(20) **(A).**
(a) $2\overset{+1}{Cu^+} \rightarrow \overset{+2}{Cu^{2+}} + \overset{0}{Cu^0}$ }Disproportionation
(b) $3\overset{+6}{MnO_4^{2-}} + 4H^+ \rightarrow 2\overset{+7}{MnO_4^-} + \overset{+4}{MnO_2} + 2H_2O$
}Disproportionation
(c) $2\overset{+7}{KMnO_4}^{-2} \xrightarrow{\Delta} K_2\overset{+6}{MnO_4} + \overset{+4}{MnO_2} + \overset{0}{O_2}$
}Not a Disproportionation
(d) $2\overset{+7}{MnO_4^-} + 3\overset{+2}{Mn^{2+}} + 2H_2O \rightarrow 5\overset{+4}{MnO_2} + 4H^+$
}Not a Disproportionation