I norganic

OXIDATION & REDUCTION

1. OXIDATION AND REDUCTION

Old Concept of Oxidation

(a) Oxidation is a chemical reaction in which oxygen is added

$2HNO_2 + O_2 \longrightarrow 2HNO_3$; $CH_3CHO + O \longrightarrow CH_3COOH$

(b) Hydrogen is removed i.e. hydrogen becomes less

$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$; $Cu + 4HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

(c) Electronegative element is added

$2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$; $2Sb + 3Cl_2 \longrightarrow 2SbCl_3$

(d) Electropositive element is removed

$2NaI + H_2O_2 \longrightarrow 2NaOH + I_2$

(e) Valency of electropositive element increases

$SnCl_2 + Cl_2 \longrightarrow SnCl_4$

Old Concept of Reduction

(a) Hydrogen is added. For example

$H_2 + Cl_2 \longrightarrow 2HCl$ $N_2 + 3H_2 \longrightarrow 2NH_3$

(b) Oxygen is lost. For example

$Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3$ $Cr_2O_3 + 2Al \longrightarrow 2Cr + Al_2O_3$

(c) Electropositive element is added. For example

$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$ $CuCl_2 + Cu \longrightarrow Cu_2Cl_2$

(d) Electronegative element is removed. For example

$2FeCl_3 + H_2 \longrightarrow 2FeCl_2 + 2HCl$ $PbS + H_2 \longrightarrow Pb + H_2S$

(e) Valency of electropositive element decreases. For example

$\underset{(Cu^{+2})}{CuSO_4} + Fe \longrightarrow FeSO_4 + \underset{(Cu^{0})}{Cu}$ $\underset{(Fe^{+3})}{FeCl_3} + H_2S \longrightarrow \underset{(Fe^{+2})}{FeCl_2} + 2HCl + S$

Modern Concept of oxidation

The reaction in which an element or an atom or an ion or molecule loses electron is called **oxidation**. de electronation is oxidation.

(a) Neutral atom : When a neutral atom loses electron, it gets converted to a positive ion.

$Na \longrightarrow Na^{+1} + e^-$ $Al \longrightarrow Al^{+3} + 3e^-$

(b) Cation : When a cation loses electron, there is an increase in its positive charge.

$Sn^{+2} \longrightarrow Sn^{+4} + 2e^-$ $Hg^{+1} \longrightarrow Hg^{+2} + e^-$

(c) Anion : When an anion loses electron equal to its negative charge, it gets converted to a neutral atom.

$2O^{-2} \longrightarrow O_2 + 4e^-$ $2N^{-3} \longrightarrow N_2 + 6e^-$

(d) Complex Anion : When a complex anion loses electron, its negative charge decreases.

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

(e) Molecule : When a molecule loses electrons, it breaks up into it constituents.

$$H_2O_2 \longrightarrow 2H^{+1} + O_2 + 2e^-$$

Therefore in oxidation reactions–

(i) Positive charge increases and negative charge decreases

(ii) Oxidation number increases

Modern Concept of Reduction

The reaction in which an element or an atom or an ion (positive or negative) or a molecule accepts electron,is called **reduction**. Electronation is reduction.

(a) Neutral Atom : When a neutral element or atom accepts electron, it get converted into an anion.

$N + 3e^- \longrightarrow N^{-3}$ $\qquad$ $S + 2e^- \longrightarrow S^{-2}$

(b) Cation : When a cation accepts electron equal to its charge, it gets converted into a neutral atom.

$Mg^{+2} + 2e^- \longrightarrow Mg^o$ $\qquad$ $Al^{+3} + 3e^- \longrightarrow Al^o$

(c) Similarly, when a cation accepts less electrons than its charge, its positive charge decreases. For example

$Cu^{+2} + e^- \longrightarrow Cu^{+1}$ $\qquad$ $Fe^{+3} + e^- \longrightarrow Fe^{+2}$

(d) Anion : When an anion accepts electron, its negative charge increases.

$MnO_4^{-1} + e^- \longrightarrow MnO_4^{-2}$ $\qquad$ $[Fe(CN)_6]^{-3} + e^- \longrightarrow [Fe(CN)_6]^{-4}$

(e) Molecule : When a molecule accepts electron, it is a reduction reaction.

$O_2 + 4e^- \longrightarrow 2O^{-2}$ $\qquad$ $I_2 + 2e^- \longrightarrow 2I^{-1}$

Therefore in reduction reactions–

(i) Positive charge decreases and negative charge increases

(ii) Oxidation number decreases

2. OXIDANTS

(i) Molecules of most electronegative elements e.g. O_2, O_3, halogens

(ii) Compounds having either of an element (under lined) in their highest oxidation state e.g. $K\underline{Mn}O_4$, $K_2\underline{Cr_2}O_7$, $H_2\underline{S}O_4$, $H\underline{N}O_3$, $\underline{Fe}Cl_3$, $\underline{Hg}Cl_2$, $KClO_3$, $Na\underline{N}O_3$ etc.

(iii) Oxides of metals and non metals e.g. MgO, CaO, CrO_3, H_2O_2, CO_2, SO_3, etc.

3. REDUCTANTS

(i) All metals e.g. Na, Al, Zn etc.

(ii) Some non metals e.g. C, S, P, H_2 etc.

(iii) Halogen acids e.g. HI, HBr, HCl.

(iv) Metallic hydrides e.g. NaH, LiH, CaH_2 etc.

(v) Compounds having either of an element (under lined) in their lowest oxidation state e.g. $\underline{Fe}Cl_2$, $\underline{Fe}SO_4$, $\underline{Hg_2}Cl_2$, $\underline{Sn}Cl_2$, $\underline{Cu_2}O$ etc.

(vi) Some organic compounds e.g. HCOOH, Aldehydes, Oxalic acid, Tartaric acid etc.

4. REDOX REACTIONS

Redox reactions are the chemical reactions which involve both oxidation as well as reduction simultaneously. In fact, oxidation and reduction go hand in hand. The redox reactions are of two types :

(i) Direct redox and (ii) Indirect redox reactions.

When chemical reactions are carried out then some of the species may lose electrons whereas some other may gain electrons. The concept of electron transfer can easily explain in the redox reactions in the case of ionic substances. However, for covalent compounds we use a new term oxidation number to explain oxidation and reduction or redox reactions. Before discussing in detail, some other terms frequently being used are:

5. SPECTATOR IONS

Species that are present in the solution but not take part in the reaction and are also omitted while writing the net ionic reaction are called spectator ions or bystander ions.

$$Zn + 2H^+ + 2Cl^- \longrightarrow Zn^{+2} + 2Cl^- + H_2$$

In this reaction ions are omitted and are called as spectator ions and appear on the reactant as well as product side.

6. TYPES OF REDOX REACTION

Autoxidation

Turpentine, Phosphorous and metals like Zn and Pb can absorb oxygen from air in the presence of water. The water oxidized to hydrogen peroxide. The phenomena of formation of H_2SO_4 by the oxidation of H_2O is known as autoxidation. $Pb + O_2 \longrightarrow PbO_2$; $PbO_2 + H_2O \longrightarrow PbO + H_2O_2$

Disproportionation

One and the same substance may act simultaneously as an oxidising and as a reducing agent. As a result a part of it gets oxidised to higher state and rest of it is reduced to a lower state of oxidation. Such as reaction, in which the substance undergoes simultaneous oxidation and reduction is called disproportionation.

Oxidation

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

Reduction

7. OXIDATION NUMBER

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

2. Oxidation number is given positive sign if electrons are lost. Oxidation number is given negative sign if electrons are gained.

3. Oxidation number represents real charge in case of ionic compounds, however, in covalent compounds it represents for imaginary charge.

4. It is the residual charge which an atom appears to have when other atom are withdrawn from the molecules as ions by containing electrons with more electronegative atoms.

The Rule for deriving Oxidation Number

Following arbitrary rules have been adopted to derive Oxidation Number of elements on the basis of periodic properties of elements.

1. In uncombined state or free state, Oxidation Number of an element is zero.

2. In combined state Oxidation Number of ....

a. .... F is always – 1.

b. .... O is –2; In peroxides (–O–O–) it is –1 and in superoxide – 1/2 .However in F_2O, it is +2.

c. .... H is 1; In ionic hydrides it is –1.

d. .... metals is always positive.

e. .... alkali metals (IA e.g. Li, Na, K, Rb, Cs, Fr) is always + 1.

f. .... alkaline earth metals (IIA e.g. Be, Mg, Ca, Sr, Ba, Ra) is always +2.

g. .... halogens in halides is always – 1.

h. .... sulphur in sulphides in always –2.

3. The algebraic sum of all the Oxidation Number of elements in a compound is equal to zero. e.g. K_2MnO_4
2 × Oxidation Number of K + Oxidation Number of Mn + 4 (Oxidation Number of O) = 0

4. The algebraic sum of all the Oxidation Numbers of elements in a radical is equal to net charge on that radical e.g.$C_2O_4^{2-}$. 2 × Oxidation Number of C + 4 (Oxidation Number of O) = – 2.

5. Oxidation Number can be zero, +ve, – ve, integer or fraction.

6. Maximum Oxidation Number of an element is (except O & F) = Group Number.
Minimum Oxidation Number of an element is (except metals) = Group Number – 8.

Note : Group number in Mendeleef's modern periodic table.

7. The most common oxidation states of some representative elements are given below.

8. Variable oxidation number is most commonly shown by transition elements as well as by p-block elements.

Transition elements : Fe (+2 & +3), Cu (+1 & +2), Mn (+7, +6, +5, +4, +3, +2, +1) etc.

p-block elements : As (+3 & +5); Sb (+3 & +5), Sn (+2 & +4) etc.

Group	Outer shell configuration	Common Oxidation Number
I gp	ns^1	0, +1
II gp	ns^2	0, +2
III gp	$ns^2\ np^1$	0, +1, +3
IV gp	$ns^2\ np^2$	0, ±1, ±2, ±3, ±4
V gp	$ns^2\ np^3$	0, ±1, ±3, +5
VI gp	$ns^2\ np^4$	0, ±2, +4, +6
VII gp	$ns^2\ np^5$	0, ±1, +3, +5, +7
Zero gp	$ns^2\ np^6$	0 (usually)

EXCEPTIONS

(i) Oxidation Number of Cl in Cl_2O is +1, because Cl acts as an electropositive element in this.

(ii) Oxidation Number of Cl in ClF_3 = +3

(iii) Oxidation Number of Cl in $KClO_3$ = +5

(iv) Oxidation Number of I in IF_7 = +7

(v) Oxidation Number of I in IF_5 = +5

Oxidation Number of Radicals

Oxidation Number of radicals is equal to charge present on them. For example,

(i) Oxidation Number of sulphite (SO_3^{-2}), sulphate (SO_4^{-2}), thiosulphate ($S_2O_3^{-2}$), oxalate ($C_2O_4^{-2}$), carbonate (CO_3^{-2}), sulphide (S^{-2}) is equal to charge (–2) present on each of them.

(ii) Oxidation Number of each of the anions, Cl^{-1}, Br^{-1}, I^{-1}, NO_3^{-1}, CN^{-1}, OH^{-1}, SCN^{-1}, CH_3COO^{-1} and HCO_3^{-1} is –1.

(iii) Oxidation Number of each of the anions. PO_4^{-3}, BO_3^{-3}, AsO_4^{-3}. (Arsenate) and AsO_3^{-3} (Arsenite) is –3.

(iv) Oxidation Number of each of the cations, CH_3^+, NH_4^+, Na^+, K^+ is +1.

(v) Oxidation Number of each of the cations, Ca^{+2}, Mg^{+2}, Sr^{+2} and Fe^{+2} is +2.

(vi) Oxidation Number of Al in $[Al(H_2O)_6]^{+3}$ is +3.

S-Element

1. S in H_2S	$2(1) + x = 0$	$+2 + x = 0$	$x = -2$
2. S in SO_2	$x + 2(-2) = 0$	$x - 4 = 0$	$x = +4$
3. S in SO_4^{-2}	$x + 4(-2) = -2$	$x - 8 = -2$	$x = +6$
4. S in SO_3^{-2}	$x + 3(-2) = -2$	$x - 6 = -2$	$x = +4$
5. S in SF_6	$x + 6(-1) = 0$	$x - 6 = 0$	$x = +6$
6. S in H_2SO_3	$2(-1) + x + 3(-2) = 0$	$+2 + x - 6 = 0$	$x = +4$
7. S in As_2S_3	$2(3) + 3x = 0$	$6 + 3x = 0$	$x = -2$

P-Element

1. Oxidation number of P in P_4 = 0

2. P in PO_4^{-3} : $x + 4(-2) = -3$ $x - 8 = -3$, $x = +5$

3. P in $NaHPO_2$: $1(1) + 1(1) + + 2(-2) = 0$ $+1 +1 + x - 4 = 0$, $x = +2$

4. P in H_3PO_3 : $3(+1) + x + 3(-2) = 0$ $+3 + x - 6 = 0$, $x = +3$

5. P in Na_2HPO_4 : $2(1) + 1(1) + x + 4(-2) = 0$ $+2 + 1 + x - 8 = 0$, $x = +5$

6. P in $Mg_2P_2O_7$: $2(2) + 2x + 7(-2) = 0$ $+4 + 2x - 14 = 0$, $2x = 10$, $x = +5$

Oxidation Number of Cr in its various compounds

1. Cr in CrO : $x - 2 = 0$, $x = +2$

2. Cr in Cr_2O_3 : $2x - 6 = 0$, $x = +3$

3. Cr in $CrSO_4$: $x - 2 = 0$, $x = +2$

4. Cr in $Cr_2(SO_4)_3$: $2x - 6 = 0$, $x = +3$

5. Cr in CrO_2Cl_2 : $2x - 6 = 0$, $x = +3$

6. Cr in $K_2Cr_2O_7$: $2 + 2x - 14 = 0$, $x = +6$

7. Cr in K_2CrO_4 : $2 + x - 8 = 0$, $x = +6$

8. Cr in $Cr_2O_7^{-2}$: $2x - 14 = -2$, $2x = 12$ $x = +6$

9. Cr in CrO_4^{-2} : $x - 8 = -2$, $x = +6$

10. Cr in Cr $(NH_3)_4SO_4$: $x - 2 = 0$, $x = +2$

(Here, Oxidation Number of NH_3 is zero)

11. Oxidation Number of Cr in $[Cr(NH_3)_4]^{+2}$: $x = +2$

12. Oxidation Number of Cr in Na_2CrO_4 : $+2 + x - 8 = 0$, $x = +6$

13. Oxidation Number of Cr in $Cr(CO)_6$: $x = 0$

(Oxidation Number of Cr = 0)

Oxidation Number of Mn in its compounds

1. Mn in MnO : $x - 2 = 0$, $x = +2$

2. Mn in Mn_2O_3 : $2x - 6 = 0$, $x = +3$

3. Mn in $MnSO_4$: $x - 2 = 0$, $x = +2$

4. Mn in $Mn_2(SO_4)_3$: $2x - 6 = 0$, $x = +3$

5. Mn in K_2MnO_4 : $+2 + x - 8 = 0$, $x = +6$

6. Mn in $KMnO_4$: $+1 + x - 8 = 0$, $x = +7$

7. Mn in $Mn(CO)_{10}$: $x + 10(0) = 0$ $x = 0$

8. Mn in MnO_4^- $x - 8 = -1$ $x = +7$

9. Mn in Mn $(C_2O_4)_2.2H_2O$: $x - 4 = 0$, $x = +4$

Oxidation state

Oxidation state of an atom is defined as oxidation number per atom.

e.g. In K_2MnO_4, Oxidation number of Mn = +6 Oxidation state of Mn = Mn^{6+}

However, for all practical purposes oxidation state is often expressed as oxidation number.

Valency and Oxidation number

Valency of an element represents the power or capacity of the element to combine with the other element.The valency of an element is numerically equal to the number of hydrogen atoms or chlorine atoms or twice the number of oxygen atoms that combine with one atom of that element. It is also equal to the number of electrons lost or accepted or shared by the atoms of an element.

In some cases (mainly in the case of electrovalent compounds), valency and oxidation number are the same but in other cases they may have different values. The difference between the two have been tabulated.

S.No.	Valency	Oxidation number (State)
1.	It is the combining capacity of the element. No plus or minus sign is attached to it.	Oxidation number is the charge (real or imaginary) present on the atom of the element when it is in combination. It may have plus or minus sign.
2.	Valency of an element is usually fixed.	Oxidation number of an element may have different values. It depends on the nature of compound in which it is present.
3.	Valency is always a whole number.	Oxidation number of the element may be a whole number or fractional.
4.	Valency of the element is never zero except in noble gases.	Oxidation number of the element may be zero.

For example, in the following compounds of carbon, the oxidation number varies from – 4 to +4 but valency of carbon is 4 in all the compounds.

Compound	CH_4	CH_3Cl	CH_2Cl_2	$CHCl_3$	CCl_4
Oxidation number of carbon	– 4	– 2	0	+ 2	+ 4

Evaluation of Oxidation Number

Determine Oxidation number of the element underlined in each of the following :

(a) **$H_2\underline{S}O_5$:**

```
        O
        ||
H-O-O-S-O-H
        ||
        O
```

$\because$ $2 \times 1 + x + 5 \times (-2) = 0$ $\therefore$ $x = +8$ (wrong)

But this can not be true as maximum oxidation number for S can not exceed +6. The exceptional value is

due to the fact that O atom in H_2SO_5 show peroxide linkage. Therefore evaluation of oxidation number should be made as :

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

(for H) (for S) (for O) (for O – O) $a = +6$

(b) **$\underline{N}H_4\underline{N}O_3$:** $2 \times x + 4 \times 1 + 3\ (-2) = 0$ $\therefore$ $x = +1$ (wrong)

No doubt NH_4NO_3 has two N atoms but one N atom has negative Oxidation Number (attached to H) and the other has +ve Oxidation Number (attached to O). Therefore, evaluation should be made separately for NH_4^+ & NO_3^-.

NH_4^+ $x + 4 \times (+1) = +1;$ $\therefore$ $x = -3$(Oxidation Number of N in NH_4^+)

NO_3^- $x + 3 \times (-2) = -1;$ $\therefore$ $x = +5$ (Oxidation Number of N in NO_3^-)

(c) **H<u>C</u>N :** The evaluation can not be made directly by using rules since we have no standard rule for oxidatio number of N and C i.e. two values are unknown. In all such cases evaluation of oxidation number should be made by indirect concept or by the original concepts of bonding.

(i) Each covalent bond contributes for one unit value for oxidation number.

(ii) Covalently bonded atom with less electronegativity acquires +ve Oxidation Number whereas other with more electronegativity acquires – ve Oxidation number.

(iii) In case of coordinate bond assign +2 value for Oxidation Number to atom from which coordinate bond is directed to other a more electronegative atom and – 2 value to more electronegative atom.

(iv) If coordinate bond is directed from more electronegative atom to less electronegative atom, then neglect contribution for coordinate bond. Thus for $H - C \equiv N$.

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

Note : $\because$ N has three covalent bonds and more electronegative than carbon.

$\therefore$ Oxidation Number of N = – 3

(d) **H – N $\equiv$ <u>C</u> :** $1 + (-3) + x = 0;$ $\therefore$ $x = +2$

[The contribution of coordinate bond is neglected because the bond is directed from more electronegative to less electronegative carbon atom.]

(e) **$\underline{Fe}_3O_4$:** $3 \times x + 4 \times (-2) = 0;$ $\therefore$ $x = +(8/3)$

or $\because$ Fe_3O_4 is a mixed oxide of FeO. Fe_2O_3

$\therefore$ Fe has two Oxidation Numbers +2 and +3.

However factually speaking Oxidation Numbers of Fe in Fe_3O_4 is an average value of these two (i.e. +2 & +3)

Average Oxidation Number $= \dfrac{1^2(+2) + 2^2(+3)}{3} = +\dfrac{8}{3}$

(f) **$\underline{Fe}SO_4\ (NH_4)_2SO_4 \cdot 6H_2O$:**

Put sum of Oxidation Numbers of $SO_4 = -2$

Sum of Oxidation Numbers in $(NH_4)_2SO_4 = 0$ [$(NH_4)_2SO_4$ is a complete molecule]

Sum of Oxidation Numbers in $H_2O = 0$ [H_2O is complete molecule]

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

(g) **$\underline{Fe}_{0.94}O$:** $x \times 0.94 + (-2) = 0;$ $x = 200/94$

(h) **$Na_2[\underline{Fe}(CN)_5NO]$:** NO in iron complexes has NO^+ nature.

Thus $2 \times 1 + x + 5 \times (-1) + 1 = 0$

(for Na) (for Fe) (for CN) (for NO); $\therefore$ $x = +2$

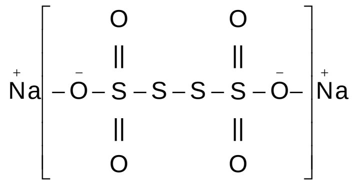
(i) **$\underline{Fe}NO(H_2O)_5SO_4$:**

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

(j) **$Na_2\underline{S}_4O_6$:**

$2 \times (+1) + 4x + 6 \times (-2) = 0; \qquad \therefore \qquad x = +5/2$

Here also this value is the average oxidation Number of S. The structure of $Na_2S_4O_6$ may be written as :



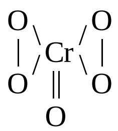
Oxidation Number of each S atom in S – S atom involved in pure covalent bond is zero.

$$\text{Average Oxidation Number } = \frac{+5+5+0+0}{4} = +\frac{5}{2}$$

(k) Dimethyl sulphoxide or $(CH_3)_2\underline{S}O$:

Oxidation Number of CH_3 = 1 : Oxidation Number of O = – 2

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

(l) $\underline{Cr}O_5$: The structure of CrO_5 has two peroxide bonds  (A butterfly structure)

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

(m) $Na_2\underline{S}_3O_6$: $\qquad 2 \times 1 + 3 \times (x) + 6 \times (-2) = 0; \qquad \therefore \qquad x = \frac{10}{3} = +3\frac{1}{3}$

8. BALANCING OF EQUATIONS :

Two methods are generally used to balance a redox equation.

By oxidation state method :

Step I & II of ion electron methods should be changed accordingly a shown

below in each case (i.e. neutral, acidic or alkaline) medium. The other steps to be followed as usual.

Example **:** $KMnO_4 + H_2C_2O_4 \longrightarrow CO_2 + K_2O + MnO + H_2O$

Step I Find the oxidation numbers of elements undergoing oxidation reduction

$Mn^{7+} \longrightarrow Mn^{2+}$ i.e. change in Oxidation Number of Mn ($+7 \longrightarrow +2$) = 5

$C_2^{3+} \longrightarrow 2C^{4+}$ i.e. change in Oxidation Number of C ($+6 \longrightarrow +8$) = 2

Step II Thus $2Mn^{7+} + 5C_2^{3+} \longrightarrow 2Mn^{2+} + 10C^{4+}$

(a) Acidic Medium : The side which has one, oxygen less is to be provided with 1 H_2O and opposite side by $2H^+$.

(b) Basic Medium : The side which has one oxygen extra is to be provided with one H_2O and opposite side by $2(OH^-)$ ions. The side which has one hydrogen extra is to be provided with $1(OH^-)$ and opposite by $1H_2O$.

Balancing of half reactions

Example 1 : $I_2 \longrightarrow IO_3^-$ (acidic medium)

Step I Balance atoms other than O & H if needed

$I_2 \longrightarrow 2IO_3^-$

Step II Balance O atoms using H^+ & H_2O as reported in step 4 of acidic medium earlier

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

Step III Balance charge by electrons

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

Example 2 : $S_2O_3^{2-} \longrightarrow SO_2$ (basic medium)

Step I $S_2O_3^{2-} \longrightarrow 2SO_2$

Step II $S_2O_3^{2-} + 2OH^- \longrightarrow 2SO_2 + H_2O$ (By H_2O & OH^-)

Step III $S_2O_3^{2-} + 2OH^- \longrightarrow 2SO_2 + H_2O + 4e^-$

Ion Electron Method :

This method involves three sets of rules depending upon the nature of equation to be balanced in neutral, acidic

or alkaline medium.

(i) Divide the overall reaction into oxidation half and reduction half reactions.

(ii) Balance the half reactions w.r.t. charges and electrons.

(iii) Equalize the electrons lost and gained by multiplying the half reactions with suitable integers. Simultaneously oxygen and Hydrogen will also be balanced.

(iv) Add the two half reactions.

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

Balancing in acidic medium

First half reaction

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

$MnO_4^- \longrightarrow Mn^{+2} + 4H_2O$

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

$5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{+2} + 4H_2O$ eq.....1

Second half reaction

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

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

Multiplying equation 2 with 5 and adding with equation 1

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

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

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

9. EQUIVALENT WEIGHT OF OXIDANTS AND REDUCTANTS

By using oxidation number, equivalent weight of oxidising and reducing substance can be determined as follows

Equivalent weight of a oxidant $= \dfrac{\text{Molecular weight of molecule or ion}}{\text{Electrons accepted by one molecule or ion}}$

$= \dfrac{\text{Molecular weight of molecule or ion}}{\text{Total change in oxidation number}}$

Equivalent weight of a reductant $= \dfrac{\text{Molecular weight of molecule or ion}}{\text{Electrons released by one molecule or ion}}$

SOLVED EXAMPLE

Ex.1 Oxidation numbers of A, B and C are +6, – 2 and –1, respectively. What will be the formula of the molecule when A, B and C associate with each other ?

[1] AB_2C_2 [2] ABC_2 [3] AB_2C [4] A_2BC

Sol. The total of positive and negative charge should be zero in the compound.

Thus, compound will be AB_2C_2 where +6 – 4 – 2 = 0

Ex.2 $3CuO + 2NH_3 \longrightarrow 3Cu + N_2 + 3H_2O$

In the above conversion, the oxidation number of nitrogen is changing in from

[1] +5 to 0 [2] 0 to +2 [3] –3 to 0 [4] –3 to –5

Sol. In $3CuO + 2NH_3 \longrightarrow 3Cu + \underset{0}{N_2} + 3H_2O$

x+ 3 = 0 x = –3 ∴ Change in 0.s = – 3 to 0

Ex.3 Oxidation numbers of the two nitrogen atoms present in ammonium nitrate are respectively ?

[1] +3 and +3 [2] 0 and 0 [3] –3 and +5 [4] –1 and –1

Sol. (i) NH_4^{+1} NO_3^{-1} Average oxidation number

X + 4 = +1 x – 6 = –1 $\frac{-3+5}{2} = +1$

x = –4 + 1, x = –3 x = +5

Ex.4 In the following reaction, $MnO_4^{-1} + 8H^+ + 5e^- \longrightarrow Mn^{+2} + 4H_2O$ how many grams of $KMnO_4$ should be taken if its 0.5 litre of 0.2 N solution is to be prepared ?

[1] 31.6 g [2] 63.2 g [3] 158.0 g [4] 94.8 g

Sol. $MnO_4^{-1} \longrightarrow Mn^{+2}$

x – 8 = –1 x = +2

x = +7

Equivalent weight = $\frac{\text{Molecular weigth}}{\text{Change in oxidation number}}$ $= \frac{158}{5}$ = 31.6 g

Weight in g = Equivalent weight × Normality × Volume = 31.6 × 0.2 × 5 = 31.6 g

Ex.5 What will be the oxidation state of copper in $YBa_2Cu_3O_7$, if oxidation state of (Y) is +3 ?

[1] 7/3 [2] 7 [3] 3 and 5 [4] none of the above

Sol. $YBa_2Cu_3O_7$

+3 + 4 + 3x – 14 = 0 3x = 7 x = 7/3

Ex.6 One mole $KMnO_4$ oxidises how many moles of ferrous oxalate ?

[1] $\frac{1}{5}$ [2] $\frac{5}{3}$ [3] $\frac{1}{3}$ [4] $\frac{2}{3}$

Sol. Reaction is

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

$C_2O_4^{-2} \rightarrow 2CO_2 + 2e] \times 5$

$5Fe^{+2} + 24H^+ + 3MnO_4^- + 5\ C_2O_4^- \rightarrow 3Mn^{+2} + 5Fe^{+3} + 10CO_2 + 12H_2O$

∵ 3 moles of $KMnO_4$ oxidises = 5 moles FeC_2O_4

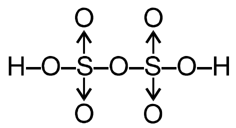
∴ 1 mole of $KMnO_4$ oxidises = $\frac{5}{3}$ moles FeC_2O_4 **Ans is 1/5**

Ex.7 What should be the oxidation number of S in $H_2S_2O_7$?

[1] +5 [2] +6 [3] +4 [4] +7

Sol. $H_2S_2O_7$

$+2 + 2x - 14 = 0$ $2x = 12$ $x = +6$



Ex.8 Oxidation number of iodine in the following reaction $IO_3^{-1} + HI \longrightarrow H_2O + I_2$

[1] increases [2] decreases

[3] increases as well as decreases [4] neither increases nor decrease

Sol. IO_3^{-1} $+ HI$ $\longrightarrow$ H_2O $+$ I_2

$x - 6 = -1$ $+ 1 + x = 0$ $x = 0$ $x = 0$

$x = +5$ $x = -1$

Oxidation number decreases from +5 to 0 and increases from –1 to 0

Ex.9 Oxidation product of Na_3AsO_3 is ?

[1] $As_2O_3^{-3}$ [2] AsO_4^{-3} [3] AsO_3 [4] AsO_2

Sol. $As_2O_3^{-3}$ AsO_4^{-3}

(Arsenite) (Arsenate)

$x - 6 = 3$ $x - 8 = -3$

$x = +3$ $x = +5$

Ex.10 One mole of X_2H_4 releases 10 moles of electrons to form a compound Y. What should be the oxidation number of X in the compound Y ?

[1] +3 [2] –3 [3] –6 [4] +1

Sol. $X_2H_4 - 10e^- \longrightarrow (X_2H_4)^{+10}$

$2x + 4 = +10$ $2x = 10 - 4 = 6$ $x = +3$

Ex.11 In the presence of humidity, SO_2

[1] loses proton [2] accepts electron [3] is an oxidant [4] is a reductant

Sol. $SO_2 + H_2O + O_2 \longrightarrow H_2SO_4$

Therefore, it changes from +4 to +6. Due to this SO_2 is a reductant.

SO_2 H_2SO_4

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

$x = +4$ $x = +6$

Ex.12 How many moles of nitrogen produced by the oxidation of one mole of hydrazine by $\frac{2}{3}$ mole bromate ion ?

[1] $\frac{1}{3}$ [2] 1 [3] 1.5 [4] $\frac{2}{3}$

Sol. The balanced equation between N_2H_4 and BrO_3^{-1} is

$3N_2H_4 + 2BrO_3^- \rightarrow 3N_2 + 2Br^- + 6H_2O$

Dividing by 3, we get : $\frac{3}{3} N_2H_4 + \frac{2}{3} BrO_3^- \rightarrow N_2 + \frac{2}{3} Br^- + 2H_2O$ **Ans is 1**

Ex.13 How many moles of $K_2Cr_2O_7$ are reduced by 1 mole of formic acid ?

[1] $\frac{1}{3}$ Mole [2] 1Mole [3] $\frac{2}{3}$ Mole [4] $\frac{5}{3}$ Mole

Sol. Equation is

$Cr_2O_7^{-2} + 8H^+ + 3HCOOH \rightarrow 2Cr^{3+} + 3CO_2 + 7H_2O$

$\because$ 3 moles of formic acid reduces = 1 mole $K_2Cr_2O_7$

$\therefore$ 1 mole of formic acid reduce = $\frac{1}{3}$ mole $K_2Cr_2O_7$ **Ans is 1/3 mole**

Ex.14 $WO_3 + 8CN^- + 2H_2O \rightarrow [W(CN)_8]^{4-} + 1/2\ O_2 + 4OH^-$ In the above process, oxidant is -

[1] WO_3 [2] CN^- [3] H_2O [4] O_2

Sol. Oxidation no. of W decreases

O.N. of W in WO_3 = +6 O.N. of W in $[W/(CN)_8]^{4-}$ = +4 **Ans is WO_3**

Ex.15 How many ml. of 0.1 M oxalic acid solution is required to reduce 0.01 mole $KMnO_4$ to MnO_2 ?

[1] 250 [2] 150 [1] 100 [4] 500

Sol. $3e + 8H^+ + MnO_4^- \rightarrow Mn^{+4} + 4H_2O$

Equivalent weight = $\frac{M}{3}$ 0.01 mole $KMnO_4$ = 0.03 equivalent $KMnO_4$

For oxalic acid : 0.1M oxalic acid = 0.2 equivalent

We have: normality = (equivalent) × $\frac{1000}{V}$

$0.2 \times 0.03 \times \frac{1000}{V}$ V = 150 ml.

Ex.16 When one mole NO_3^- is converted into 1 mole NO_2, 0.5 mole. N_2 and 0.5 mole N_2O respectively. It accepts x, y and z mole of electrons –x, y and z are respectively.

[1] 1, 5, 4 [2] 1, 2, 3 [3] 2, 1, 3 [4] 2, 3, 4

Sol. The equation are :

$NO_3^- + 2H^+ + e \rightarrow NO_2 + H_2O$

$NO_3^- + 6H^+ + 5e \rightarrow 0.5N_2 + 3H_2O$

$NO_3^- + 5H^+ + 4e \rightarrow 0.5N_2O + 2.5H_2O$

$\therefore$ x, y and z respectively are 1, 5 and 4.

Ex.17 Calculate the equivalent weight of potassium permanganate ($KMnO_4$) in (i) neutral medium (ii) acidic medium (iii) alkaline medium, by oxidation number method.

Sol. (i) $Mn^{+7} + 3e \rightarrow Mn^{+4}$; Eq. wt. = M/3

(ii) $Mn^{+7} + 5e \rightarrow Mn^{+2}$; Eq. wt. = M/5

(iii) $Mn^{+7} + 1e \rightarrow Mn^{+6}$; Eq. wt. = M/1

Ex.18 An element A in a compound ABD has an oxidation no. A^{-n}. It is oxidised by $Cr_2O_7^{-2}$ in acid medium. In an experiment 1.68×10^{-3} mole of $K_2Cr_2O_7$ was required for 3.26×10^{-3} mole of the compound ABD. Calculate new oxidation state of A.

Sol. $A^{-n} \longrightarrow A^{+a} + (a + n)e$

$6e + Cr_2^{+6} \longrightarrow 2Cr^{+3}$

$\therefore$ Meq. of A^{-n} = Meq. of $Cr_2O_7^{-2}$ or $3.26 \times 10^{-3} \times (a + n) = 1.68 \times 10^{-3} \times 6$

$\therefore$ $a + n = 3$ or $a =$ **3 – n**

Ex.19 Find out the value of n in $MnO_4^- + 8H^+ + ne \rightarrow Mn^{+2} + 4H_2O$

Sol. $\therefore$ Total charge on L.H.S. = Total charge on R.H.S.

$-1 + 8 - (-n) = +2;$ $\therefore n = 5$

Ex.20 In the reaction $8\ Al + 3\ Fe_3O_4 \rightarrow 4\ Al_2O_3 + 9\ Fe$

(a) Which element is oxidised or reduced ?

(b) Total number of electrons transferred during the change.

Sol. $8\ Al^0 \rightarrow 4Al_2^{3+} + 24e$

$24e + 3Fe_3^{(8/3)+} \rightarrow 9Fe^0$

or $8Al^0 + 3\ Fe_3^{(8/3)+} \rightarrow 4\ Al_2^{3+} + 9Fe$

Reductant is Al i.e. **Al is oxidised**

Oxidant is Fe_3O_4 or $Fe^{(8/3)+}$ i.e. **$Fe^{(8/3)+}$ is reduced**

Number of electrons used during redox change = **24**

Ex.21 A student unsuccessfully tried to balance the following equation :

$Cr_2O_7^{2-} + Fe^{3+} + H^+ \rightarrow Cr^{3+} + Fe^{2+} + H_2O$. Why could not student balance the equation?

Sol. Both parts are reduction part i.e. Cr^{+6} as well as Fe^{3+} both are reduced without a reductant which is not possible.

Ex.22 Six moles of Cl_2 undergo a loss and gain of 10 moles of electrons to form two oxidation state of Cl.

Write down the two half reactions & find out the oxidation number of each Cl atom involved.

Sol. $6Cl_2 \rightarrow 2\ Cl^{5+} + 10\ Cl^-$

+ 5; –1;

Ex.23 Reaction between 1 mole of $HgCl_2$ and 1 mole of $SnCl_2$ occurs as follows. $2\ HgCl_2 + SnCl_2 \rightarrow SnCl_4 + Hg_2Cl_2$. Which of the following ions will be there after completion of the reaction?

[1] $Hg^{+1}, Sn^{+2}, Sn^{+4}$ [2] Hg^{+2}, Sn^{+2} [3] Sn^{+2}, Sn^{+4} [4] $Hg^{+2}, Sn^{+2}, Sn^{+4}$

Sol. According to the reaction, 2 mole $HgCl_2$ reacts with 1 mole $SnCl_2$. Therefore, 1 mole $HgCl_2$ will react with 1/2 mole $SnCl_2$ & 1/2 mole $SnCl_2$ will be left. Thus, Sn^{+4}, Hg^{+1} and Sn^{+2} ions will remain in the solution.

EXERCISE - 1

OXIDATION REDUCTION DEFINITION

1. Reduction is defined as :

[1] Increase in positive valency [2] Gain of electrons

[3] Loss of protons [4] Decrease in negative valency

2. $Co(s) + Cu^{2+}(aq) \rightarrow Co^{2+}(aq) + Cu(s)$. The above reaction is :

[1] Oxidation reaction [2] Reduction reaction [3] Redox reaction [4] None of these

3. Which of the following reactions depict the oxidising behavior of H_2SO_4 :

[1] $2PCl_5 + H_2SO_4 \rightarrow 2POCl_3 + 2HCl + SO_2Cl_2$ [2] $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

[3] $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$ [4] $2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$

4. In $C + H_2O \rightarrow CO + H_2$, H_2O acts as :

[1] Oxidising agent [2] Reducing agent [3] Both [4] None

5. Reducing agent is that :

[1] Which takes electrons [2] Which takes protons

[3] Which donates electrons [4] Which donates protons

6. HBr and HI reduce sulphuric acid. HCl can reduce $KMnO_4$ and HF can reduce :

[1] H_2SO_4 [2] $KMnO_4$ [3] $K_2Cr_2O_7$ [4] None of these

7. The compound which gives oxygen on moderate heating is :

[1] Ferric oxide [2] Zinc oxide [3] Mercuric oxide [4] Aluminium oxide

8. In a reaction between zinc and iodine in which zinc iodide is formed, what is being oxidised :

[1] Zinc ions [2] Iodide ions [3] Zinc atom [4] Iodine

9. In the following reactions : $4P + 3KOH + 3H_2O \rightarrow 3KH_2PO_2 + PH_3$

[1] Only phosphorus is oxidized [2] Only phosphorus is reduced

[3] Phosphorus is both oxidized and reduced [4] Phosphorus is neither oxidized nor reduced

10. The reaction of $Zn^{++} + 2e^- \rightarrow Zn$ is an example of :

[1] Oxidatio [2] Reduction [3] Redox reaction [4] None

11. In the reaction $3Cl_2 + 6OH^- \rightarrow 5Cl^- + ClO_3^- + 3H_2O$ chlorine is :

[1] Oxidised [2] Reduced

[3] Oxidised as well as reduced [4] Neither oxidised nor reduced

12. In the reaction $3Br_2 + 6CO_3^{2-} + 3H_2O \rightarrow 5Br^- + BrO_3^- + 6HCO_3^-$

[1] Bromine is oxidised and carbonate is reduced [2] Bromine is both reduced and oxidised

[3] Bromine is neither reduced nor oxidised [4] Bromine is reduced and water is oxidised

13. A gas X bleaches a flower by reduction and another gas Y by oxidation these gases are , respectively

[1] NH_3 & SO_3 [2] NO_2 & N_2O_5 [3] SO_2 & Cl_2 [4] SO_2 & PCl_3

14. What will happen when copper rod is dipped in aluminium nitrate solution, if the electropositive properties are as follows : Al < Zn > Cu > Ag

[1] Aluminium will get deposited on the rod [2] Colour of the solution will becomes blue

[3] Copper aluminium alloy will be formed [4] No reaction will occur

15. For the reaction : $4Fe + 3O_2 \rightarrow 4Fe^{3+} + 6O^{2-}$ which of the following is a wrong statement ?

[1] It is an example of redox reaction [2] Metallic iron reduces to Fe^{3+}

[3] Fe is oxidised [4] Metallic iron is a reducing agent

16. In the reaction

$MnO_4^- + NO_2^- \rightarrow NO_3^- + Mn^{2+}$

one mole of MnO_4^- oxidises ________ moles of NO_2^-

[1] 5 [2] 5/2 [3] 3 [4] 3/2

17. In the following equation $ClO_3^- + 6H^+ + X \rightarrow Cl^- + 3H_2O$, then X is

[1] O [2] $6e^-$ [3] O_2 [4] $5e^-$

18. Which one of the following compounds can act as an oxidising as well as reducing agent -

[1] $KMnO_4$ [2] H_2O_2 [3] BaO [4] $K_2Cr_2O_7$

19. When acidic solution of ferrous ammonium sulphate is treated with potassium permanganate solution then the ion which is oxidised is -

[1] MnO_4^- [2] NH_4^+ [3] Fe^{2+} [4] SO_4^{2-}

20. The violent reaction between sodium and water is an example of -

[1] Reduction [2] Oxidation [3] Redox reaction [4] Neutralization

21. In the formation of $Pb(NO_3)_2$ form PbO_2 -

[1] PbO_2 is oxidised [2] PbO_2 is reduced

[3] PbO_2 is both oxidised and reduced. [4] PbO_2 is neither oxidised nor reduced

22. Which of the following s an example of reduction -

[1] $CuO \rightarrow Cu_2O$ [2] $[Fe(CN)_6]^{4-} \rightarrow [Fe(CN)_6]^{3-}$

[3] $KI \rightarrow I_2$ [4] $H_2S \rightarrow S$

23. Reaction $[Ag(NH_3)_2]^+ + 2H^+ \rightarrow Ag^+ + 2NH_4^+$ is an example of -

[1] Oxidation [2] Reduction

[3] Neither oxidation nor reduction [4] Oxidation and reduction both

24. Which of the following reactions involves neither oxidation nor reduction -

[1] $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$ [2] $Cr^- \rightarrow CrCl_3$ [3] $VO^{2+} \rightarrow V_2O_2$ [4] $2S_2O_3^{2-} \rightarrow S_4O_6^{2-}$

25. What would happen when a small quantity of H_2O_2 is added to a solution of $FeSO_4$ -

[1] Colour disappears [2] H_2 is evolved

[3] An electron is added to Fe^{++} [4] An electron is lost by Fe^{++}

26. The reaction $2TiCl_3 \rightarrow TiCl_2 + TiCl_4$ example of -

[1] dissociation [2] disproportionation [3] reversible reaction [4] exothermic reaction

27. The anodic reaction in the electrolysis of the aqueous solution of NaCl is -

[1] Oxidation of chloride ion [2] Evolution of oxygen

[3] reduction of chloride ion [4] Oxidation of sodium ion.

28. In the reaction -

$2FeCl_3 + H_2S \rightarrow 2FeCl_2 + 2HCl + S$

[1] $FeCl_3$ is used as an oxidant. [2] $FeCl_3$ and H_2S both are oxidised.

[3] $FeCl_3$ is oxidised and H_2S is reduced. [4] H_2S is used as an oxidant.

RULES OF OXIDATION NUMBER AND OXIDATION NUMBER

29. A compound contains atoms X, Y and Z the oxidation number of X is + 2, Y is + 5 and Z is – 2 therefore a possible formula of the compound is :

[1] XYZ_2 [2] $X_2(YZ_3)_2$ [3] $X_3(YZ_4)_2$ [4] $X_3(Y_4Z)_2$

30. The atomic number of an element which shows the oxidation state of + 3 is :

[1] 13 [2] 32 [3] 33 [4] 17

31. Which of the following is the correct oxidation number of phosphorus in $Mg_2P_2O_7$:

[1] – 3 [2] + 2 [3] + 5 [4] + 3

32. Oxidation number of sulphur in Na_2SO_4 is :

[1] – 2 [2] + 6 [3] + 2 [4] – 6

33. Oxidation state of O_2 in H_2O_2 is :

[1] – 2 [2] – 1 [3] + 1 [4] + 2

34. If three electrons are lost by a metal ion M^{3+}, its final oxidation number should be :

[1] 0 [2] + 6 [3] + 2 [4] + 4

35. Oxidation number of Fe in $K_3[Fe(CN)_6]$ is :

[1] + 2 [2] + 3 [3] + 1 [4] + 4

36. Oxidation number of sulphur in S_2Cl_2 is :

[1] + 1 [2] 0 [3] – 1 [4] + 6

37. Oxidation number of sulphur in $S_2O_2^{2-}$ is :

[1] – 2 [2] + 1 [3] + 6 [4] 0

38. Oxidation number of nitrogen in NH_3 is :

[1] – 3 [2] + 3 [3] 0 [4] + 5

39. The oxidation number of nitrogen in NH_2OH is :

[1] + 1 [2] – 1 [3] – 3 [4] – 2

40. Oxidation number of P in KH_2PO_2 is :

[1] +1 [2] 6 [3] 4 [4] 7

41. In the compounds $KMnO_4$ and $K_2Cr_2O_7$, the highest oxidation state is of the element :

[1] Potassium [2] Manganese [3] Chromium [4] Oxygen

42. The normal oxidation state of an element is – 2. The number of electrons in its outermost shell will be

[1] 4 [2] 2 [3] 6 [4] 8

43. Oxidation number of Ni in $Ni(CO)_4$ is :

[1] 0 [2] 4 [3] 8 [4] 2

44. The oxidation number of nitrogen in NH_4NO_3 is :

[1] + 3 [2] + 5 [3] – 3 and + 5 [4] + 3 and + 5

45. Which of the following halogens always shows only one oxidation state ?

[1] Cl [2] F [3] Br [4] I

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

[1] ICl [2] ClO_3^- [3] ClF_3 [4] $HClO_4$

47. The oxidation number of cobalt in $[Co(CN)_6]^{3-}$ is -

[1] + 3 [2] - 3 [3] + 6 [4] - 6

48. In which of the following compound oxidation number of iron is not +3

[1] Fe_3O_4 [2] Fe_2O_3 [3] $FeCl_3$ [4] $FePO_4$

49. The oxidation number of Mn in MnC_2O_4 is -

[1] + 3 [2] + 8/3 [3] + 1 [4] +2

50. The correct oxidation number of phosphorus in magnesium pyrophosphate $[Mg_2P_2O_7]$ is -

[1] + 2 [2] + 3 [3] - 3 [4] + 5

51. Oxidation number of sulphur in H_2SO_5 is -

[1] + 2 [2] + 4 [3] + 8 [4] + 6

52. In which of the following compound, iodine is in its highest oxidation state -

[1] KI [2] KIO_4 [3] KI_3 [4] IF_5

53. Oxidation number of chlorine in Hypochlorous acid is–

[1] –1 [2] zero [3] + 1 [4] + 2

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

[1] $Fe_2(CO)_9$ [2] $Ni(CO)_4$ [3] $Fe_3(CO)_9$ [4] All of the above

55. The oxidation state of phosphorus is + 3 in -

[1] Orthophosphorous acid [2] Orthophosphoric acid

[3] Pyrophosphoric acid [4] Metaphosphoric acid

56. Which of the following is a true statement -

[1] Oxidation state of oxygen in HOF is zero. [2] Oxidation state of fluorine in HOF is – 1.

[3] Oxidation state of chlorine in HOCl is + 1. [4] All of the above.

57. The following reaction is used in the extraction of chromium from its ore

$2Fe_2O_3.Cr_2O_3 + 4Na_2CO_3 + 3O_2 \rightarrow 2Fe_2O_3 + 4Na_2CrO_4 + 4CO_2$

What is true about the oxidation states of the substance in the reaction -

[1] Chromium is oxidised from + 3 to + 6 oxidation state.

[2] Iron is reduced from + 3 to + 2 oxidation state.

[3] Carbon is oxidised from + 3 to + 4 oxidation state

[4] There is no change in the oxidation states of the substances.

58. Oxidation state of nitrogen is incorrectly given for -

Compounds	Oxidation states	Compounds	Oxidation states
[1] $[Co(NH_3)_5Cl]Cl_2$	– 3	[2] NH_2OH	– 1
[3] $(N_2H_5)_2SO_4$	+2	[4] Mg_3N_2	– 3

59. Out of the following acids which has different oxidation state of phosphorus as compared to others -

[1] Phosphorous acid [2] Orthophosphoric acid [3] Metaphosphoric acid [4] Pyrophosphoric acid

60. The brown ring complex compound is formulated as $[Fe(H_2O)_5NO^+]SO_4$. The oxidation state of iron is -

[1] 1 [2] 2 [3] 3 [4] zero

61. When $KMnO_4$ is reduced with oxalic acid in acidic solution, the oxidation number of Mn changes from -

[1] 7 to 4 [2] 6 to 4 [3] 7 to 2 [4] 4 to 2

62. The oxidation number of each sulphur in $Na_2S_4O_6$ is -

[1] 2.5 [2] 2 and 3 (two S have + 2 and the other two have + 3)
[3] 2 and 4 (three S have + 2 and one S has + 4) [4] 5 and 0 (two S have + 5 and the other S have 0)

63. In a triatomic molecule the oxidation states of atoms A, B and C are + 6, + 1 and – 2 respectively. The molecular formula of the compound will be -

[1] B_2AC_4 [2] $B_2A_2C_7$ [3] Both of the above. [4] None of the above

64. Which of the following statements is not correct -

[1] Two mole of electrons are used in the reduction of MnO_4^- to MnO_3^-

[2] Three electrons per chromium atom are used in the reduction of dichromate by Fe (II)

[3] The oxidation state of oxygen is $-\frac{1}{2}$ in potassium superoxide.

[4] The oxidation number increases in the process of reduction.

BALANCING EQUATION

65. In acidic medium, reaction : $MnO_4^- \rightleftharpoons Mn^{2+}$ is an example of :

[1] Oxidation by three electrons [2] Reduction by three electrons

[3] Oxidation by five electrons [4] Reduction by five electrons

66. In which of the following reaction there is no change in valency

[1] $4KClO_3 \rightarrow 3KClO_4 + KCl$ [2] $SO_2 + 2H_2S \rightarrow 2H_2O + 3S$

[3] $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$ [4] $2BaO + O_2 \rightarrow 2BaO_2$

67. In the reaction : $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2SO_4$

[1] Valency of barium increases [2] Valency of barium is not changed

[3] Valency of barium becomes zero [4] Valency of barium decreases

68. A balanced half reaction for the unbalanced whole reaction

$CrO_4^{-2} + SO_3^{-2} \rightarrow CrO_2^- + SO_4^{-2} + OH^-$

[1] $2CrO_4^{-2} + 8H_2O + 6e \rightarrow 2CrO_2^- + 4H_2O + 8OH^-$

[2] $2CrO_4^{-2} + 8H_2O \rightarrow CrO_2^- + 4H_2O + 8OH^-$

[3] $CrO_4^{-2} + H_2O \rightarrow CrO_2^- + H_2O + OH^-$

[4] $3CrO_4^{-2} + 4H_2O + 6e \rightarrow 2CrO_2^- + 8OH^-$

69. Choose the set of the coefficient that correctly balance the equation

$xCr_2O_7^{-2} + yH^+ + ze \rightarrow aCr^{+3} + bH_2O$

	x	y	z	a	b		x	y	z	a	b
[1]	2	14	6	2	7	[2]	1	14	6	2	7
[3]	2	7	6	2	7	[4]	2	7	6	1	7

70. $8Al + 3Fe_3O_4 \rightarrow 4Al_2O_3 + 9Fe$

In the reaction how many total electrons will be transferred

[1] 12 [2] 24 [3] 20 [4] 14

71. How many protons will be added to the right to balance the process $MnO_4^- \rightarrow Mn^{+2}$

[1] 0 [2] 8 [3] 5 [4] 2

72. Reactions (a and b)

(a) $Fe^{+3} + H_2O_2 \rightarrow Fe^{+2} + O_2$ (b) $Cr(OH)_2 + I_2 \rightarrow Cr(OH)_3 + 2I^-$

Should be balanced in acidic or basic medium

[1] a (acidic), b (basic) [2] a (acidic), b (acidic) [3] a (basic), b (basic) [4] a (basic), b (acidic)

73. In the reaction, $A^{-n}{}_2 + xe \rightarrow A^{-n}{}_1$

Here x will be

[1] $n_1 + n_2$ [2] $n_2 - n_1$ [3] $n_1 - n_2$ [4] n_1n_2

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

The coefficient of the reactants in the balanced state of the equation are

[1] 1, 3 [2] 1, 4 [3] 2, 2 [4] 2, 4

EQUIVALENT MASS

75. In acidic medium equivalent weight of $K_2Cr_2O_7$ (molecular weight = M) is :

[1] M / 3 [2] M / 4 [3] M / 6 [4] M / 2

76. In the following reaction

$As_2S_5 + NO_3^- \rightarrow AsO_4^{3-} + SO_4^{2-} + NO_2$

The equivalent weight of As_2S_5 is

[1] M/8 [2] M/6 [3] M/40 [4] M/30

77. In a reaction the equivalent weight of $KMnO_4$ becomes one third of its molecular weight. The oxidation state of Mn in the final product is

[1] + 6 [2] + 4 [3] + 3 [4] + 2

78. The equivalent weight of reducing agent in the reaction

$2[Fe(CN)_6]^{3-} + 2OH^- + H_2O_2 \rightarrow 2[Fe(CN)_6]^{4-} + 2H_2O + O$ -

[1] 17 [2] 212 [3] 16 [4] 6/8

79. In a redox reaction $K_2Cr_2O_7$ changes to $Cr_2(SO_4)_3$. If the molecular weight of $K_2Cr_2O_7$ is M and equivalent weight E then -

[1] M = 3E [2] M = 6E [3] E = 2M [4] E = 6M

80. Fe_3O_4 is oxidised to Fe_2O_3. If the molecular weight of Fe_3O_4 is M and equivalent weight E then -

[1] E = M [2] $E = \frac{M}{3}$ [3] $E = \frac{2}{3}M$ [4] $E = \frac{3}{2}M$

Answer Key

Qus.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	2	3	4	1	3	4	3	2	3	2	3	2	3	4	2	2	2	2	3	3
Qus.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	1	1	3	1	4	2	1	1	3	1	3	2	2	2	2	1	2	1	2	1
Qus.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	2	3	1	3	2	3	1	1	4	4	4	2	3	4	1	4	1	3	1	1
Qus.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	3	4	3	4	4	3	2	1	2	2	1	1	3	2	3	3	2	1	2	1

EXERCISE - 2

1. Oxidation state of Cr in $Cr(CO)_6$ is - **[AIIMS–93]**

[1] 0 [2] + 2 [3] - 2 [4] + 6

2. Oxidation number of Pt in $[Pt(C_2H_4)Cl_3]^-$ is - **[MLNR–93]**

[1] + 1 [2] + 2 [3] + 3 [4] + 4

3. The brown ring complex is formulated as $[Fe(H_2O)_5NO]SO_4$. The oxidation state of iron is -

[1] + 1 [2] + 2 [3] + 3 [4] 0 **[MPPMT–93]**

4. The oxidation number of sulphur in S_8, S_2F_2, H_2S respectively, are - **[IIT–1999]**

[1] 0, + 1 and - 2 [2] + 2, +1 and - 2 [3] 0, + 1 and + 2 [4] - 2, + 1 and - 2

5. Which of the following is not a reducing agent - **[DCE–2000]**

[1] SO_2 [2] H_2O_2 [3] CO_2 [4] NO_2

6. Equivalent mass of oxidising agent in the reaction, **[DCE–2000]**

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

[1] 32 [2] 64 [3] 16 [4] 8

7. A, B and C are three element forming a part of compound in oxidation states of + 2, + 5 and - 2 respectively. What could be the compound - **[CPMT–2000]**

[1] $A_2(BC)_2$ [2] $A_2(BC_4)_3$ [3] $A_3(BC_4)_2$ [4] ABC

8. On reduction of $KMnO_4$ by oxalic acid in acidic medium, the oxidation number of Mn changes. What is the magnitude of this change - **[MPPMT–2000]**

[1] From 7 to 2 [2] From 6 to 2 [3] From 5 to 2 [4] From 7 to 4

9. The oxidation number of iron in Fe_3O_4 is - **[CEET–2000]**

[1] + 2 [2] + 3 [3] 8/3 [4] 2/3

10. What is oxidation number of Fe in $Fe(CO)_5$ - **[CPMT–2000]**

[1] Zero [2] 5 [3] – 5 [4] + 3

11. In H_2O_2, the oxidation state of oxygen is - **[CPMT–2000]**

[1] – 2 [2] – 1 [3] 0 [4] 4

12. In the balanced equation - **[CPMT–2000]**

$5H_2O_2 + XClO_2 + 2OH^- \rightarrow XCl^- + YO_2 + 6H_2O$

The reaction is balanced if

[1] X = 5, Y = 2 [2] X = 2, Y = 5 [3] X = 4, Y = 10 [4] X = 5, Y = 5

13. Best way to prevent rusting of iron is by - **[DPMT–2000]**

[1] making iron cathode [2] putting it in saline water

[3] both of these [4] none of these

14. Amongst the following, identify the species with an atom in + 6 oxidation state - **[IIT–2000]**

[1] MnO_4^- [2] $Cr(CN)_6^{3-}$ [3] NiF_6^{2-} [4] CrO_2Cl_2

15. HNO_3 acts as - **[MANIPAL–2001]**

[1] acid [2] oxidising agent [3] reducing agent [4] Both (a) and (b)

16. The reaction, **[IIT–2001]**

$3ClO^{-}(aq) \rightarrow ClO_3^{-}(aq) + 2Cl^{-}(aq)$

is an example of-

[1] Oxidation reaction [2] Reduction reaction

[3] Disproportionation reaction [4] Decomposition reaction.

17. In the standardization of $Na_2S_2O_3$ using $K_2Cr_2O_7$ by using iodometry, the equivalent weight of $K_2Cr_2O_7$ is -

[1] (molecular weight) / 2 [2] (molecular weight) / 6 **[IIT–2001]**

[3] (molecular weight) / 3 [4] same as molecular weight.

18. The oxidation number of sulphur in $Na_2S_4O_6$ is - **[RPMT–2002]**

[1] 1.5 [2] 2.5 [3] 3 [4] 2

19. Which of the following is a redox reaction - **[AIEEE–2002]**

[1] $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$ [2] $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$

[3] $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + NH_4OH$ [4] $Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$

20. When $KMnO_4$ acts as an oxidising agent and ultimately forms MnO_4^{2-}, MnO_2, Mn_2O_3 and Mn^{2+} then the number of electrons transferred in each case respectively is - **[AIEEE–2002]**

[1] 4, 3, 1, 5 [2] 1, 5, 3, 7 [3] 1, 3, 4, 5 [4] 3, 5, 7, 1

21. The oxidation state of Fe in $K_4[Fe(CN)_6]$ is - **[CET–2002]**

[1] + 2 [2] + 6 [3] + 3 [4] + 4

22. Oxidation number of S in $H_2S_2O_8$ is - **[CET–2002]**

[1] + 2 [2] + 4 [3] + 6 [4] + 7

23. Which reaction is not feasible - **[CPMT–2002]**

[1] $2KI + Br_2 \rightarrow 2KBr + I_2$ [2] $2KBr + I_2 \rightarrow 2KI + Br_2$

[3] $2KBr + Cl_2 \rightarrow 2KCl + Br_2$ [4] $2H_2O + 2F_2 \rightarrow 4HF + O_2$

24. The oxidation state of Cr in $[Cr(NH_3)_4Cl_2]^+$ is **[AIEEE–2005]**

[1] 0 [2] +1 [3] +2 [4] +3

25. The oxidation state of chromium in the final product formed by the reaction between KI and acidified potassium dichromate solution is **[AIEEE–2005]**

[1] +3 [2] +2 [3] +6 [4] +4

Answer Key

Qus.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	1	2	1	1	3	3	2	1	3	1	2	2	1	4	4	3	2	2	4	3
Qus.	**21**	**22**	**23**	**24**	**25**															
Ans.	1	3	2	4	1															