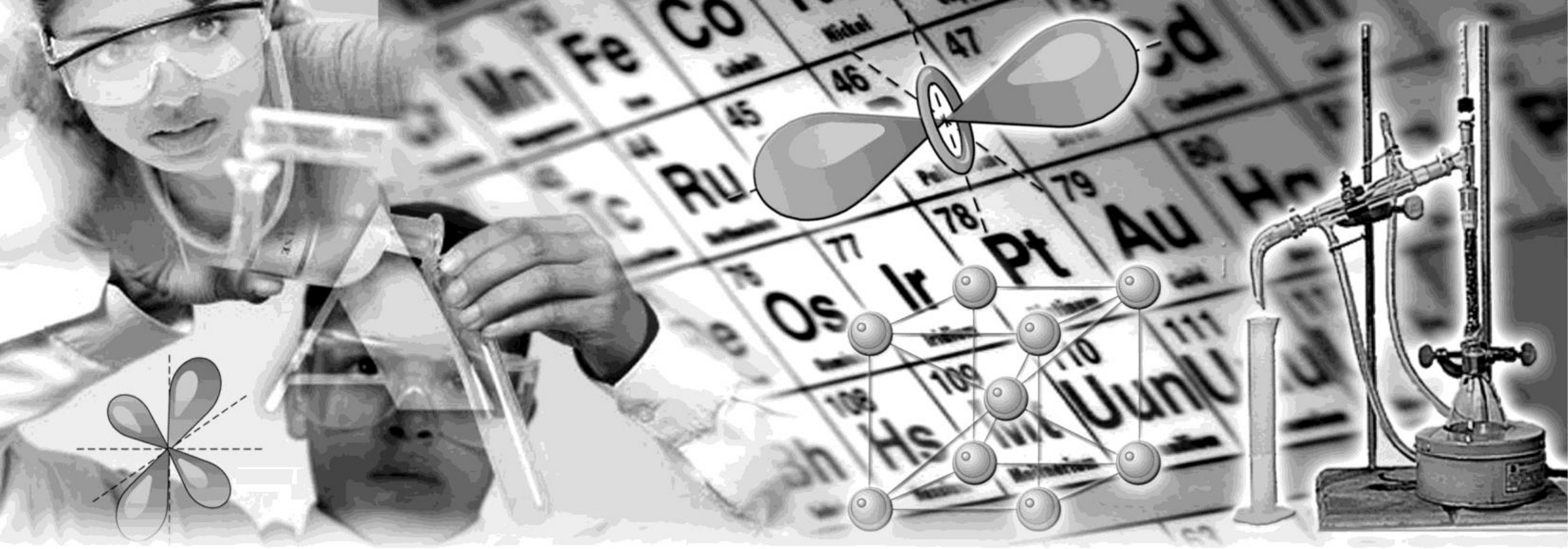


Chapter 13
Redox Reactions

Chemical reactions involve transfer of electrons from one chemical substance to another. These electron – transfer reactions are termed as oxidation-reduction or redox-reactions.

Molecular and Ionic equations

(1) **Molecular equations :** When the reactants and products involved in a chemical change are written in molecular forms in the chemical equation, it is termed as molecular equation.

Example : $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

In above example the reactants and products have been written in molecular forms, thus the equation is termed as ***molecular equation.***

(2) **Ionic equations :** When the reactants and products involved in a chemical change are ionic compounds, these will be present in the form of ions in the solution. The chemical change is written in ionic forms in chemical equation, it is termed as ionic equation. *Example*,

$$MnO_2 + 4H^+ + 4Cl^- \rightarrow Mn^{2+} + 2Cl^- + 2H_2O + Cl_2$$

In above example the reactants and products have been written in ionic forms, thus the equation is termed as ***ionic equation.***

(3) **Spectator ions :** In ionic equations, the ions which do not undergo any change and equal in number in both reactants and products are termed as spectator ions and are not included in the final balanced equations. *Example,*

$Zn + 2H^+ + 2Cl^- \rightarrow Zn^{2+} + H_2 + 2Cl^-$ (*Ionic equation*)

$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$ (*Final ionic equation*)

In above example, the Cl^- ions are the ***spectator ions*** and hence are not included in the final ionic balanced equation.

Oxidation-reduction and Redox reactions

(1) **Oxidation :** Oxidation is a process which involves; addition of oxygen, removal of hydrogen, addition of non-metal, removal of metal, Increase in +*ve* valency, loss of electrons and increase in oxidation number.

(i) ***Addition of oxygen*** **:** $2Mg + O_2 \rightarrow 2MgO$

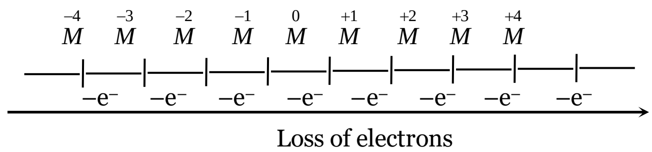
(ii) ***Removal of hydrogen*** **:** $H_2S + Cl_2 \rightarrow 2HCl + S$

(iii) ***Addition of Non-metal*** **:** $Fe + S \rightarrow FeS$

(iv) ***Removal of metal*** **:** $2KI + H_2O_2 \rightarrow 2KOH + I_2$

(v) ***Increase in +ve valency*** **:** $Fe^{2+} \rightarrow Fe^{3+} + e^-$

(vi) ***Loss of electrons*** (also known as de-electronation)



(a) $H^0 \rightarrow H^+ + e^-$ (*Formation of proton*)

(b) $MnO_4^{2-} \rightarrow MnO_4^- + e^-$ (*De-electronation of* MnO_4^{2-})

(c) $2Fe^0 \rightarrow 2Fe^{3+} + 6e^-$ (*De-electronation of iron*)

(vii) ***Increase in oxidation number***

(a) $Mg^0 \rightarrow Mg^{2+}$ (*From* 0 *to* +2)

(b) $[Fe^{+2}(CN)_6]^{4-} \rightarrow [Fe^{+3}(CN)_6]^{3-}$ (From +2 *to* +3)

(c) $2Cl^- \rightarrow Cl_2^0$ (From –1 *to* 0)

(2) **Reduction :** Reduction is just reverse of oxidation. Reduction is a process which involves; removal of oxygen, addition of hydrogen, removal of non-metal, addition of metal, decrease in *+ve* valency, gain of electrons and decrease in oxidation number.

(i) ***Removal of oxygen :*** $CuO + C \rightarrow Cu + CO$

(ii) ***Addition of hydrogen :*** $Cl_2 + H_2 \rightarrow 2HCl$

(iii) ***Removal of non-metal***

$2HgCl_2 + SnCl_2 \rightarrow Hg_2Cl_2 + SnCl_4$

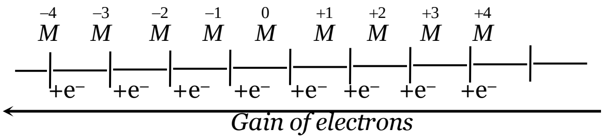
(iv) ***Addition of metal :*** $HgCl_2 + Hg \rightarrow Hg_2Cl_2$

(v) ***Decrease in +ve valency***

(a) $Fe^{3+} \rightarrow Fe^{2+}$ (*+ve valency decreases*)

(b) $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}$ (*–ve valency increases*)

(vi) ***Gain of electrons*** (also known as electronation)



(a) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(S)$ (*Electronation of* Zn^{2+})

(b) $Pb^{2+} + 2e^- \rightarrow Pb^0$ (*Electronation of* Pb^{2+})

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

(*Electronation of* $[Fe(CN)_6]^{3-}$)

(vii) ***Decrease in oxidation number***

(a) $Mg^{2+} \rightarrow Mg^0$ (*From* +2 *to* 0)

(b) $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}$ (*From* +3 *to* +2)

(c) $Cl_2^0 \rightarrow 2Cl^-$ (*From* 0 *to* –1)

(3) **Redox-reactions**

(i) An overall reaction in which oxidation and reduction takes place simultaneously is called ***redox*** or ***oxidation-reduction reaction.*** These reactions involve transfer of electrons from one atom to another. Thus every redox reaction is made up of two ***half reactions***; One half reaction represents the oxidation and the other half reaction represents the reduction.

(ii) ***Types of redox reaction***

(a) *Direct redox reaction* : The reactions in which oxidation and reduction takes place in the same vessel are called direct redox reactions.

(b) *Indirect redox reaction* : The reactions in which oxidation and reduction takes place in different vessels are called indirect redox reactions. Indirect redox reactions are the basis of electro-chemical cells.

(c) *Intermolecular redox reactions* : In which one substance is oxidised while the other is reduced.

For example, $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

Here, *Al* is oxidised to Al_2O_3 while Fe_2O_3 is reduced to Fe.

(d) *Intramolecular redox reactions* : In which one element of a compound is oxidised while the other is reduced.

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

Here, Cl^{+5} in $KClO_3$ is reduced to Cl^{-1} in *KCl* while O^{2-} in $KClO_3$ is oxidised to O_2^0.

Oxidising and Reducing agents

(1) **Definition :** The substance (atom, ion or molecule) that gains electrons and is thereby reduced to a low valency state is called an ***oxidising agent***, while the substance that loses electrons and is thereby oxidised to a higher valency state is called a ***reducing agent.***

Or

An ***oxidising agent*** is a substance, the oxidation number of whose atom or atoms decreases while a *reducing agent* is a substance the oxidation number of whose atom increases.

(2) **Important oxidising agents**

(i) Molecules made up of electronegative elements.

Example : O_2, O_3 and X_2 (halogens).

(ii) Compounds containing an element which is in the highest oxidation state.

Example : $KMnO_4, K_2Cr_2O_7, Na_2Cr_2O_7, CrO_3, H_2SO_4,$ $HNO_3, NaNO_3, FeCl_3, HgCl_2, KClO_4, SO_3, CO_2, H_2O_2$ etc.

(iii) Oxides of elements, $MgO, CuO, CrO_3, CO_2, P_4O_{10}$, etc.

(iv) Fluorine is the strongest oxidising agent.

(3) **Important reducing agents**

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

(ii) A few non-metals *e.g. C,* H_2, *S* etc.

(iii) Hydracids : *HCl, HBr, HI,* H_2S etc.

(iv) A few compounds containing anelement in the lower oxidation state (ous).

Example : $FeCl_2$, $FeSO_4$, $SnCl_2$, Hg_2Cl_2, Cu_2O etc.

(v) Metallic hydrides *e.g. NaH, LiH* etc.

(vi) Organic compounds like *HCOOH* and $(COOH)_2$ and their salts, aldehydes, alkanes etc.

(vii) Lithium is the strongest reducing agent in solution.

(viii) Cesium is the strongest reducing agent in absence of water. Other reducing agents are $Na_2S_2O_3$ and *KI.*

(ix) Hypo prefix indicates that central atom of compound has the minimum oxidation state so it will act as a reducing agent.

Example : H_3PO_2 (hypophosphorous acid).

(4) **Substances which act as oxidising as well as reducing agents**

Examples : $H_2O_2, SO_2, H_2SO_3, HNO_2, NaNO_2, Na_2SO_3, O_3$ etc.

(5) **Tips 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.

(6) **Equivalent weight of oxidising and reducing agents**

Equivalent weight of a substance (oxidant or reductant) is equal to molecular weight divided by number of electrons lost or gained by one molecule of the substance in a redox reaction.

$$\text{Eq. wt. of } \mathbf{O.\ A.} = \frac{\text{Molecular weight}}{\text{No. of electrons gained by one molecule}}$$

$$= \frac{\text{Molecular weight}}{\text{Change in O. N. per mole}}$$

$$\text{Eq. wt. of } \mathbf{R.\ A.} = \frac{\text{Molecular weight}}{\text{No. of electrons lost by one molecule}}$$

$$= \frac{\text{Molecular weight}}{\text{Change in O. N. per mole}}$$

Table : 13.1 Equivalent weight of few oxidising/reducing agents

Agents	O. N.	Product	O. N.	Change in O. N. per atom	Total Change in O. N. per mole	Eq. wt.
$Cr_2O_7^{2-}$	+ 6	Cr^{3+}	+ 3	3	3 × 2 = 6	Mol. wt./6
$C_2O_4^{2-}$	+ 3	CO_2	+ 4	1	1 × 2 = 2	Mol. wt./2
$S_2O_3^{2-}$	+ 2	$S_4O_6^{2-}$	+ 2.5	0.5	0.5 × 2 = 1	Mol. wt./1
H_2O_2	– 1	H_2O	– 2	1	1 × 2 = 2	Mol. wt./2
H_2O_2	– 1	O_2	0	1	1 × 2 = 2	Mol. wt./2
MnO_4^- **(Acidic medium)**	+ 7	Mn^{2+}	+ 2	5	5 × 1 = 5	Mol. wt./5
MnO_4^- **(Neutral medium)**	+ 7	MnO_2	+ 4	3	3 × 1 = 3	Mol. wt./3
MnO_4^- **(Alkaline medium)**	+ 7	MnO_4^{2-}	+ 6	1	1 × 1 = 1	Mol. wt./1

Oxidation number or Oxidation state

(1) **Definition :** Charge on an atom produced by donating or accepting electrons is called ***oxidation number*** or ***oxidation state.*** It is the number of effective charges on an atom.

(2) **Valency and oxidation number :** Valency and oxidation number concepts are different. 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. Points of difference between the two have been tabulated below

Oxidation number	Valency
O.N. 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.	It is the combining capacity of the element. No plus or minus sign is attached to it.
O.N. of an element may have different values. It depends on the nature of compound in which it is present.	Valency of an element is usually fixed.
O.N. of the element may be a whole number or fractional.	Valency is always a whole number.
O.N. of the element may be zero.	Valency of the element is never zero except of noble gases.

(3) **Oxidation number and Nomenclature**

(i) When an element forms two monoatomic cations (representing different oxidation states), the two ions are distinguished by using the ending-*ous* and *ic. The suffix – ous is used for the cation with lower oxidation state and the suffix – ic is used for the cation with higher oxidation state.*

For example : Cu^+ (oxidation number +1) cuprous

Cu^{2+} (oxidation number +2) cupric

(ii) ***Albert Stock*** proposed a new system known as *Stock system*. In this system, Roman numeral written in parentheses immediately after the name of the element indicates the oxidation states. For example,

Cu_2O	Copper (I) oxide	SnO	Tin (II) oxide
$FeCl_2$	Iron (II) chloride	Mn_2O_7	Manganess (VII) oxide
$K_2Cr_2O_7$	Potassium dichromate (VI)	Na_2CrO_4	Sodium chromate (VI)

V_2O_5	Vanadium (V) oxide	CuO	Copper (II) oxide
SnO_2	Tin (IV) oxide	$FeCl_3$	Iron (III) chloride

(4) **Rules for the determination of oxidation number of an atom :** The following rules are followed in ascertaining the oxidation number of an atom,

(i) If there is a covalent bond between two same atoms then oxidation numbers of these two atoms will be zero. Bonded electrons are symmetrically distributed between two atoms. Bonded atoms do not acquire any charge. So oxidation numbers of these two atoms are zero.

$A : A$ or $A - A \rightarrow A^* + A^*$

For example, Oxidation number of Cl in Cl_2, O in O_2 and N in N_2 is zero.

(ii) If covalent bond is between two different atoms then electrons are counted towards more electronegative atom. Thus oxidation number of more electronegative atom is negative and oxidation number of less electronegative atom is positive. Total number of charges on any element depends on number of bonds.

$A - B \longrightarrow A^+ + B^- :$

$A - B \longrightarrow A^{+2} + :B^{-2}:$

The oxidation number of less electronegative element (A) is + 1 and + 2 respectively.

(iii) If there is a coordinate bond between two atoms then oxidation number of donor atom will be + 2 and of acceptor atom will be – 2.

$A \rightarrow B \longrightarrow A^{2+} + :B^{-2}:$

(iv) The oxidation number of all the atoms of different elements in their respective elementary states is taken to be zero. For example, in $N_2, Cl_2, H_2, P_4, S_8, O_2, Br_2, Na, Fe, Ag$ etc. the oxidation number of each atom is zero.

(v) The oxidation number of a monoatomic ion is the same as the charge on it. For example, oxidation numbers of Na^+, Mg^{2+} and Al^{3+} ions are + 1, + 2 and + 3 respectively while those of Cl^-, S^{2-} and N^{3-} ions are –1, –2 and –3 respectively.

(vi) The oxidation number of hydrogen is + 1 when combined with non-metals and is –1 when combined with active metals called metal hydrides such as LiH, KH, MgH_2, CaH_2 etc.

(vii) The oxidation number of oxygen is – 2 in most of its compounds, except in peroxides like H_2O_2, BaO_2 etc. where it is –1. Another interesting exception is found in the compound OF_2 (*oxygen difluoride*) where the oxidation number of oxygen is + 2. This is due to the fact that fluorine being the most electronegative element known has always an oxidation number of –1.

(viii) In compounds formed by union of metals with non-metals, the metal atoms will have positive oxidation numbers and the non-metals will have negative oxidation numbers.

For example,

(a) The oxidation number of alkali metals (*Li, Na, K* etc.) is always +1 and those of alkaline earth metals (*Be, Mg, Ca* etc) is + 2.

(b) The oxidation number of halogens (*F, Cl, Br, I*) is always –1 in metal halides such as KF, $AlCl_3$, $MgBr_2$, CdI_2. etc.

(ix) In compounds formed by the union of different elements, the more electronegative atom will have negative oxidation number whereas the less electronegative atom will have positive oxidation number.

For example,

(a) N is given an oxidation number of –3 when it is bonded to less electronegative atom as in NH_3 and NI_3, but is given an oxidation number of + 3 when it is bonded to more electronegative atoms as in NCl_3.

(b) Since fluorine is the most electronegative element known so its oxidation number is always –1 in its compounds i.e. oxides, interhalogen compounds etc.

(c) In interhalogen compounds of *Cl, Br,* and *I;* the more electronegative of the two halogens gets the oxidation number of –1. For example, in $BrCl_3$, the oxidation number of Cl is –1 while that of Br is +3.

(x) For neutral molecule, the sum of the oxidation numbers of all the atoms is equal to zero. For example, in NH_3 the sum of the oxidation numbers of nitrogen atom and 3 hydrogen atoms is equal to zero. For a complex ion, the sum of the oxidation numbers of all the atoms is equal to charge on the ion. For example, in SO_4^{2-} ion, the sum of the oxidation numbers of sulphur atom and 4 oxygen atoms must be equal to –2.

(xi) It may be noted that oxidation number is also frequently called as oxidation state. For example, in H_2O, the oxidation state of hydrogen is +1 and the oxidation state of oxygen is – 2. This means that oxidation number gives the oxidation state of an element in a compound.

(xii) In the case of representative elements, the highest oxidation number of an element is the same as its group number while highest negative oxidation number is equal to (8 – Group number) with negative sign with a few exceptions. The most common oxidation states of the representative elements are shown in the following table,

Group	Outer shell configuration	Common oxidation numbers (states) except zero in free state
I A	ns^1	+1
II A	ns^2	+2
III A	ns^2np^1	+3, +1
IV A	ns^2np^2	+4,+3,+2,+1, –1, –2, –3, –4
V A	ns^2np^3	+5,+3,+1, –1, –3
VI A	ns^2np^4	+6,+4,+2,–2
VII A	ns^2np^5	+7,+5,+3, +1, –1

(xiii) Transition metals exhibit a large number of oxidation states due to involvement of $(n-1)$ d electron besides ns electron.

(xiv) Oxidation number of a metal in carbonyl complex is always zero.

Example : Ni has zero oxidation state in $[Ni(CO)_4]$.

(xv) Those compounds which have only C, H and O the oxidation number of carbon can be calculated by following formula,

$$\text{Oxidation number of } 'C' = \frac{(n_O \times 2 - n_H)}{n_C}$$

Where, n_O is the number of oxygen atom, n_H is the number of hydrogen atom, n_C is the number of carbon atom.

For example, (a) $\overset{*}{C}H_3OH$; $n_H = 4, n_C = 1, n_O = 1$

Oxidation number of 'C' = $\frac{(1 \times 2 - 4)}{1} = -2$

(b) $H\overset{*}{C}OOH$; $n_H = 2$, $n_O = 2, n_c = 1$

Oxidation number of carbon = $\frac{(2 \times 2 - 2)}{1} = +2$

(5) **Procedure for calculation of oxidation number :** By applying the above rules, we can calculate the oxidation numbers of elements in the molecules/ions by the following steps.

(i) Write down the formula of the given molecule/ion leaving some space between the atoms.

(ii) Write oxidation number on the top of each atom. In case of the atom whose oxidation number has to be calculated write x.

(iii) Beneath the formula, write down the total oxidation numbers of each element. For this purpose, multiply the oxidation numbers of each atom with the number of atoms of that kind in the molecule/ion. Write the product in a bracket.

(iv) Equate the sum of the oxidation numbers to zero for neutral molecule and equal to charge on the ion.

(v) Solve for the value of x.

Table : 13.2 Oxidation number of some elements in compounds, ions or chemical species

Element	Oxidation Number	Compounds, ions or chemical species
Sulphur (S)	− 2	H_2S, ZnS, $NaHS$, $(SnS_3)^{2-}$, BaS, CS_2
	0	S, S_4, S_8, SCN^-
	+ 1	S_2F_2, S_2Cl_2
	+ 4	SO_2, H_2SO_3, $(SO_3)^{2-}$, $SOCl_2$, $NaHSO_3$, $Ca[HSO_3]_2$, $[HSO_3]^-$, SF_4
	+ 6	H_2SO_4, $(SO_4)^{2-}$, $[HSO_4]^-$, $BaSO_4$, $KHSO_4$, SO_3, SF_6, $H_2S_2O_7$, $(S_2O_7)^{2-}$
Nitrogen (N)	− 3	NH_3, $(NH_4)^+$, AlN, Mg_3N_2, $(N)^{3-}$, Ca_3N_2, CN^-
	− 2	N_2H_4, $(N_2H_5)^+$
	− 1	NH_2OH
	−1/3	NaN_3, N_3H
	0	N_2
	+ 1	N_2O
	+ 2	NO
	+ 3	HNO_2, $(NO_2)^-$, $NaNO_2$, N_2O_3, NF_3
	+ 4	NO_2
	+ 5	HNO_3, $(NO_3)^-$, KNO_3, N_2O_5
Chlorine (Cl)	− 1	HCl, $NaCl$, $CaCl_2$, $AlCl_3$, ICl, ICl_5, $SOCl_2$, CrO_2Cl_2, KCl, K_2PtCl_6, $HAuCl_4$, CCl_4
	0	Cl, Cl_2
	+ 1	$HOCl$, $NaOCl$, $(OCl)^-$, Cl_2O
	+ 3	$KClO_2$, $(ClO_2)^-$, $HClO_2$
	+ 4	ClO_2
	+ 5	$(ClO_3)^-$, $KClO_3$, $NaClO_3$, $HClO_3$
	+ 7	$HClO_4$, Cl_2O_7, $KClO_4$, $(ClO_4)^-$

Hydrogen (*H*)	− 1	NaH, CaH_2, $LiAlH_4$, LiH
	+ 1	NH_3, PH_3, HF
Phosphorus (*P*)	− 3	PH_3, $(PH_4)^+$, Ca_3P_2
	0	P_4
	+ 1	H_3PO_2, KH_2PO_2, $BaH_4P_2O_4$
	+ 3	PI_3, PBr_3, PCl_3, P_2O_3, H_3PO_3
	+ 5	$(PO_4)^{3-}$, H_3PO_4, $Ca_3(PO_4)_2$, $H_4P_2O_7$, P_4O_{10}, PCl_5, $(P_2O_7)^{4-}$, $Mg_2P_2O_7$, *ATP*
Oxygen (*O*)	− 2	H_2O, PbO_2, $(CO_3)^{2-}$, $(PO_4)^{2-}$, SO_2, $(C_2O_4)^{2-}$, $HOCl$, $(OH)^-$, $(O)^{2-}$
	− 1	Na_2O_2, BaO_2, H_2O_2, $(O_2)^{2-}$, **Peroxides**
	− 1/2	KO_2
	0	O, O_2, O_3
	+ 1	O_2F_2
	+ 2	OF_2
Carbon (*C*)	− 4	CH_4
	− 3	C_2H_6
	− 2	CH_3Cl, C_2H_4
	− 1	CaC_2, C_2H_2
	0	Diamond, Graphite, $C_6H_{12}O_6$, $C_2H_4O_2$, $HCHO$, CH_2Cl_2
	+ 2	CO, $CHCl_3$, HCN
	+ 3	$H_2C_2O_4$, $(C_2O_4)^{2-}$
	+ 4	CO_2, H_2CO_3, $(HCO_3)^-$, CCl_4, Na_2CO_3, Ca_2CO_3, CS_2, CF_4, $(CO_3)^{-2}$
Chromium (*Cr*)	+ 3	$Cr_2(SO_4)_3$, $CrCl_3$, Cr_2O_3, $[Cr(H_2O)_4Cl_3]$
	+ 6	K_2CrO_4, $(CrO_4)^{2-}$, $K_2Cr_2O_7$, $(Cr_2O_7)^{2-}$, $KCrO_3Cl$, CrO_2Cl_2, $Na_2Cr_3O_{10}$, CrO_3
Manganese (*Mn*)	+ 2	MnO, $MnSO_4$, $MnCl_2$, $Mn(OH)_2$
	+ 8/3	Mn_3O_4
	+ 3	$Mn(OH)_3$
	+ 4	MnO_2, K_2MnO_3
	+ 6	K_2MnO_4, $(MnO_4)^{2-}$
	+ 7	$KMnO_4$, $(MnO_4)^-$, $HMnO_4$
Silicon (*Si*)	− 4	SiH_4, Mg_2Si
	+ 4	SiO_2, K_2SiO_3, $SiCl_4$
Iron (*Fe*)	$+\frac{8}{3}$	Fe_3O_4
	+ 2	$FeSO_4.(NH_4)_2SO_4$ (Ferrous ammonium sulphate), $K_4Fe(CN)_6$, $FeCl_2$
	+ 3	$K_3[Fe(CN)_6]$, $FeCl_3$
Iodine (*I*)	+ 7	$H_4IO_6^-$, KIO_4
Osmium (*Os*)	+ 8	OsO_4

Xenon(*Xe*)	+ 6	***Xe***O_3, ***Xe***F_6

(6) **Exceptional cases of evaluation of oxidation numbers :** The rules described earlier are usually helpful in determination of the oxidation number of a specific atom in simple molecules but these rules fail in following cases. In these cases, the oxidation numbers are evaluated using the concepts of chemical bonding involved.

Type I. In molecules containing peroxide linkage in addition to element-oxygen bonds. For example,

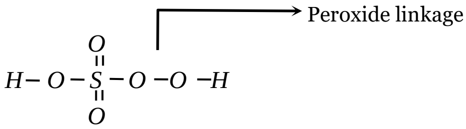
(i) ***Oxidation number of S in*** H_2SO_5

(Permonosulphuric acid or Caro's acid)

By usual method; H_2SO_5

$2\times 1+x+5\times(-2)=0$ or $x=+8$

But this cannot be true as maximum oxidation number for S cannot exceed + 6. Since S has only 6 electrons in its valence shell. This exceptional value is due to the fact that two oxygen atoms in H_2SO_5 shows peroxide linkage as shown below,



Therefore the evaluation of o.n. of sulphur here should be made as follows,

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

(for H) (for S) (for O) (for O–O)

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

(ii) **Oxidation number of *S* in** $H_2S_2O_8$

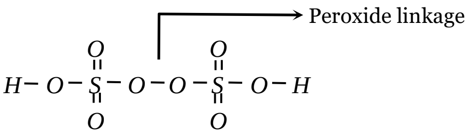
(Peroxidisulphuric acid or Marshall's acid)

By usual method ; $H_2S_2O_8$

$1\times 2+2x+8(-2)=0$

$2x=+16-2=14$ or $x=+7$

Similarly Caro's acid, Marshall's acid also has a peroxide linkage so that in which S shows +6 oxidation state.



Therefore the evaluation of oxidation state of sulphur should be made as follow,

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

(for H) (for S) (for O) (for O–O)

or $2+2x-12-2=0$ or $x=+6$.

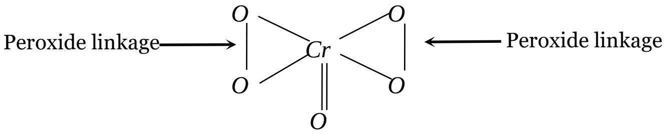
(iii) **Oxidation number of *Cr* in** CrO_5

(Blue perchromate)

By usual method CrO_5 ; $x-10=0$ or $x=+10$

This cannot be true as maximum O. N. of Cr cannot be more than + 6. Since *Cr* has only five electrons in 3d orbitals and one electron in 4s orbital. This exceptional value is due to the fact that four oxygen atoms in CrO_5 are in peroxide linkage.

The chemical structure of CrO_5 is



Therefore, the evaluation of o.n. of *Cr* should be made as follows

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

(for *Cr*) (for *O*) (for *O–O*)

or $x-2-4=0$ or $x=+6$.

Type II. In molecules containing covalent and coordinate bonds, following rules are used for evaluating the oxidation numbers of atoms.

(i) For each covalent bond between dissimilar atoms the less electronegative element is assigned the oxidation number of + 1 while the atom of the more electronegative element is assigned the oxidation number of –1.

(ii) In case of a coordinate-covalent bond between similar or dissimilar atoms but the donor atom is less electronegative than the acceptor atom, an oxidation number of +2 is assigned to the donor atom and an oxidation number of –2 is assigned to the acceptor atom.

Conversely, if the donor atom is more electronegative than the acceptor atom, the contribution of the coordinate bond is neglected. *Examples,*

(a) *Oxidation number of C in* $HC\equiv N$ *and* $HN\overset{\rightarrow}{=}C$

The evaluation of oxidation number of *C* cannot be made directly by usual rules since no standard rule exists for oxidation numbers of *N* and *C*.

In such cases, evaluation of oxidation number should be made using indirect concept or by the original concepts of chemical bonding.

(b) *Oxidation number of carbon in* $H-N\overset{\rightarrow}{=}C$

The contribution of coordinate bond is neglected since the bond is directed from a more electronegative *N* atom (donor) to a less electronegative carbon atom (acceptor).

Therefore the oxidation number of *N* in $HN\overset{\rightarrow}{=}C$ remains – 3 as it has three covalent bonds.

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

(for *H*) (for *N*) (for *C*)

or $1+x-3=0$ or $x=+2$.

(c) *Oxidation number of carbon in* $HC\equiv N$

In $HC\equiv N$, *N* is more electronegative than carbon, each bond gives an oxidation number of –1 to *N*. There are three covalent bonds, the oxidation number of *N* in $HC\equiv N$ is taken as – 3

Now $HC\equiv N$ $\therefore$ $+1+x-3=0\Rightarrow x=+2$

Type III. In a molecule containing two or more atoms of same or different elements in different oxidation states.

(i) ***Oxidation number of S in $Na_2S_2O_3$***

By usual method $Na_2S_2O_3$

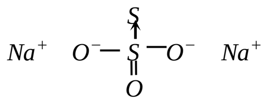
$\therefore\ 2 \times (+1) + 2 \times x + 3\,(-2) = 0$ or $2 + 2x - 6 = 0$

or $x = 2$.

But this is unacceptable as the two sulphur atoms in $Na_2S_2O_3$ cannot have the same oxidation number because on treatment with dil. H_2SO_4, one sulphur atom is precipitated while the other is oxidised to SO_2.

$$Na_2S_2O_3 + H_2SO_4 \rightarrow Na_2SO_4 + SO_2 + S + H_2O$$

In this case, the oxidation number of sulphur is evaluated from concepts of chemical bonding. The chemical structure of $Na_2S_2O_3$ is



Due to the presence of a co-ordinate bond between two sulphur atoms, the acceptor sulphur atom has oxidation number of – 2 whereas the other S atom gets oxidation number of + 2.

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

(for Na) (for O) (for S) (for coordinated S)

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

Thus two sulphur atoms in $Na_2S_2O_3$ have oxidation number of – 2 and +6.

(ii) ***Oxidation number of chlorine in $CaOCl_2$***

(bleaching powder)

In bleaching powder, $Ca(OCl)Cl$, the two Cl atoms are in different oxidation states *i.e.*, one Cl^- having oxidation number of –1 and the other as OCl^- having oxidation number of +1.

(iii) ***Oxidation number of N in NH_4NO_3***

By usual method $N_2H_4O_3$; $2x + 4 \times (+1) + 3 \times (-1) = 0$

$2x + 4 - 3 = 0$ or $2x = +1$ *(wrong)*

No doubt NH_4NO_3 has two nitrogen atoms but one N has negative oxidation number (attached to H) and the other has positive oxidation number (attached to O). Hence the evaluation should be made separately for NH_4^+ and NO_3^-

NH_4^+ $x + 4 \times (+1) = +1$ or $x = -3$

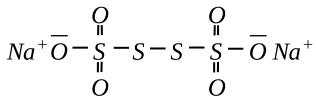
NO_3^- $x + 3\ (-2) = -1$ or $x = +5$.

(iv) ***Oxidation number of Fe in Fe_3O_4***

In Fe_3O_4, Fe atoms are in two different oxidation states. Fe_3O_4 can be considered as an equimolar mixture of FeO [iron (II) oxide] and Fe_2O_3 [iron (III) oxide]. Thus in one molecule of Fe_3O_4, two Fe atoms are in + 3 oxidation state and one Fe atom is in + 2 oxidation state.

(v) ***Oxidation number of S in sodium tetrathionate*** ($Na_2S_4O_6$)

Its structure can be represented as follows,



The two S-atoms which are linked to each other have oxidation number zero. The oxidation number of other S-atoms can be calculated as follows

Let oxidation number of $S = x$.

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

(for S) (for S–S) (for O)

$x = +5$.

Balancing of oxidation-reduction reactions

Though there are a number of methods for balancing oxidation – reduction reactions, two methods are very important. These are,

(1) Oxidation number method

(2) Ion – electron method

(1) **Oxidation number method** : The method for balancing redox reactions by oxidation number change method was developed by ***Johnson.*** In a balanced redox reaction, total increase in oxidation number must be equal to the total decrease in oxidation number. This equivalence provides the basis for balancing redox reactions. This method is applicable to both molecular and ionic equations. The general procedure involves the following steps,

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

(ii) Assign oxidation numbers to the atoms in the equation and find out which atoms are undergoing oxidation and reduction. Write separate equations for the atoms undergoing oxidation and reduction.

(iii) Find the change in oxidation number in each equation. Make the change equal in both the equations by multiplying with suitable integers. Add both the equations.

(iv) Complete the balancing by inspection. First balance those substances which have undergone change in oxidation number and then other atoms except hydrogen and oxygen. Finally balance hydrogen and oxygen by putting H_2O molecules wherever needed.

The final balanced equation should be checked to ensure that there are as many atoms of each element on the right as there are on the left.

(v) In ionic equations the net charges on both sides of the equation must be exactly the same. Use H^+ ion/ions in acidic reactions and OH^- ion/ions in basic reactions to balance the charge and number of hydrogen and oxygen atoms.

The following example illustrate the above rules,

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

(*Skeleton equation*)

Step**:** II Writing the oxidation number of all the atoms.

$$\overset{0}{Cu} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} \rightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2 + \overset{+4}{N}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

Step**:** III Change in oxidation number has occurred in copper and nitrogen.

$$\overset{0}{Cu} \rightarrow \overset{+2}{Cu}(NO_3)_2 \qquad ......(i)$$

$$H\overset{+5}{N}O_3 \rightarrow \overset{+4}{N}O_2 \qquad ......(ii)$$

Increase in oxidation number of copper = 2 units per molecule *Cu*

Decrease in oxidation number of nitrogen = 1 unit per molecule HNO_3

Step: IV To make increase and decrease equal, equation (ii) is multiplied by 2.

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

Step**:** V Balancing nitrate ions, hydrogen and oxygen, the following equation is obtained.

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

This is the balanced equation.

(2) **Ion-electron method** (half reaction method)

Jette and **LaMev** developed the method for balancing redox-reactions by ion electron method in 1927. It involves the following steps

(i) Write down the redox reaction in ionic form.

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

(iii) Balance each half reaction for the number of atoms of each element. For this purpose,

(a) Balance the atoms other than *H* and *O* for each half reaction using simple multiples.

(b) Add water molecules to the side deficient in oxygen and H^+ to the side deficient in hydrogen. This is done in acidic or neutral solutions.

(c) In alkaline solution, for each excess of oxygen, add one water molecule to the same side and $2OH^-$ ions to the other side. If hydrogen is still unbalanced, add one OH^- ion for each excess hydrogen on the same side and one water molecule to the other side.

(iv) Add electrons to the side deficient in electrons as to equalise the charge on both sides.

(v) Multiply one or both the half reactions by a suitable number so that number of electrons become equal in both the equations.

(vi) Add the two balanced half reactions and cancel any term common to both sides.

The following example illustrate the above rules

Step**:** I $I_2 + OH^- \rightarrow IO_3^- + I^- + H_2O$ *(Ionic equation)*

Step**:** II Splitting into two half reactions,

$$I_2 + OH^- \rightarrow IO_3^- + H_2O \ ; \quad I_2 \rightarrow I^-$$

(Oxidation half reaction) *(Reduction half reaction)*

Step**:** III Adding OH^- ions, $I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O$

Step**:** IV Adding electrons to the sides deficient in electrons, (*Si*)

$$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^- \ ; \ I_2 + 2e^- \rightarrow 2I^-$$

Step**:** V Balancing electrons in both the half reactions.

$$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^- \ ; \ 5[I_2 + 2e^- \rightarrow 2I^-]$$

Step**:** VI Adding both the half reactions.

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

Dividing by 2, $3I_2 + 6OH^- \rightarrow IO_3^- + 5I^- + 3H_2O$

Autoxidation

(1) Turpentine and numerous other olefinic compounds, phosphorus and certain metals like *Zn* and *Pb* can absorb oxygen from the air in presence of water. The water is oxidised to hydrogen peroxide. This phenomenon of formation of H_2O_2 by the oxidation of H_2O is known as ***autoxidation.*** The substance such as turpentine or phosphorus or lead which can activate the oxygen is called ***activator.*** The activator is supposed to first combine with oxygen to form an addition compound, which acts as an ***autoxidator*** and reacts with water or some other acceptor so as to oxidise the latter. For example;

$$\underset{\text{(activator)}}{Pb} + O_2 \rightarrow \underset{\text{(autoxidator)}}{PbO_2} \ ; \ PbO_2 + \underset{\text{(acceptor)}}{H_2O} \rightarrow PbO + H_2O_2$$

(2) The turpentine or other unsaturated compounds which act as activators are supposed to take up oxygen molecule at the double bond position to form unstable peroxide called ***moloxide,*** which then gives up the oxygen to water molecule or any other acceptor.

$$RCH = CHR + O_2 \rightarrow \underset{O-O}{RHC-CHR}$$

$$\underset{O-O}{RHC-CHR} + 2H_2O \rightarrow RCH = CHR + 2H_2O_2$$

$$2KI + H_2O_2 \rightarrow 2KOH + I_2$$

The evolution of iodine from *KI* solution in presence of turpentine can be confirmed with starch solution which turns blue.

(3) The concept of autoxidation help to explain the phenomenon of induced oxidation. Na_2SO_3 solution is oxidised by air but Na_3AsO_3 solution is not oxidised by air. If a mixture of both is taken, it is observed both are oxidised. This is induced oxidation.

$$Na_2SO_3 + O_2 \rightarrow \underset{\text{Moloxide}}{Na_2SO_5}$$

$$Na_2SO_5 + Na_3AsO_3 \rightarrow Na_3AsO_4 + Na_2SO_4$$

$$Na_2SO_3 + Na_3AsO_3 + O_2 \rightarrow Na_2SO_4 + Na_3AsO_4$$

Disproportionation

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

Following are the some examples of *disproportionation*,

(1) $\overset{-1}{H_2O_2} + H_2\overset{-1}{O_2} = H_2\overset{-2}{O} + \overset{0}{O_2}$ (−1 → 0: increase; −1 → −2: decrease)

(2) $4K\overset{+5}{Cl}O_3 \rightarrow 3K\overset{+7}{Cl}O_4 + K\overset{-1}{Cl}$ (+5 → −1: decrease; +5 → +7: increase)

(3) $4\overset{0}{P} + 3NaOH + 3H_2O \rightarrow 3NaH_2\overset{+1}{P}O_2 + \overset{-3}{P}H_3$ (0 → −3: decrease; 0 → +1: increase)

(4) $3\overset{0}{Cl_2} + \underset{(conc.)}{6NaOH} \xrightarrow{hot} 5Na\overset{-1}{Cl} + Na\overset{+5}{Cl}O_3 + 3H_2O$ (0 → −1: decrease; 0 → +5: increase)

Tips & Tricks

- If an element is in its highest possible oxidation state in a compound, it can act as an oxidising agent. for example, $KMnO_4$, $K_2Cr_2O_7$, HNO_3, H_2SO_4, $HClO_4$ etc.
- If an element is in its lowest oxidation state in a compound, it can act as a reducing agent. For example, H_2S, $H_2C_2O_4$, $FeSO_4$, $Na_2S_2O_3$, SO_2, $SnCl_2$, many metals etc.
- The strength of oxyacids of chlorine decrease in the order. $HClO_4 > HClO_3 > HClO_2 > HClO$
- If highly electronegative element is in its highest oxidation state in a compound that compound can act as powerful oxidant. For example, $KClO_4$, $KClO_3$, $KBrO_3$, KIO_3 etc.
- If an element is in intermediate oxidation state in a compound, it can act as both oxidising & reducing agent. For example, H_2O_2, H_2SO_3, HNO_3, SO_2 etc.

Ordinary Thinking

Objective Questions

Oxidation, Reduction

1. H_2O_2 reduces MnO_4^- ion to **[KCET (Med.) 2000]**

(a) Mn^+ (b) Mn^{2+}
(c) Mn^{3+} (d) Mn^-

2. When a sulphur atom becomes a sulphide ion **[AMU 1999]**

(a) There is no change in the composition of atom
(b) It gains two electrons
(c) The mass number changes
(d) None of these

3. The ultimate products of oxidation of most of hydrogen and carbon in food stuffs are **[DCE 2001]**

(a) H_2O alone (b) CO_2 alone
(c) H_2O and CO_2 (d) None of these

4. When P reacts with caustic soda, the products are PH_3 and NaH_2PO_2. This reaction is an example of **[IIT 1980; Kurukshetra CEE 1993; CPMT 1997]**

(a) Oxidation
(b) Reduction
(c) Oxidation and reduction (Redox)
(d) Neutralization

5. Which one of the following does not get oxidised by bromine water **[MP PET/PMT 1988]**

(a) Fe^{+2} to Fe^{+3} (b) Cu^+ to Cu^{+2}
(c) Mn^{+2} to MnO_4^- (d) Sn^{+2} to Sn^{+4}

6. In the reaction $H_2S + NO_2 \rightarrow H_2O + NO + S$. H_2S is

(a) Oxidised (b) Reduced
(c) Precipitated (d) None of these

7. The conversion of PbO_2 to $Pb(NO_3)_2$ is

(a) Oxidation
(b) Reduction
(c) Neither oxidation nor reduction
(d) Both oxidation and reduction

8. In the course of a chemical reaction an oxidant **[MP PMT 1986]**

(a) Loses electrons
(b) Gains electrons
(c) Both loses and gains electron
(d) Electron change takes place

9. $2CuI \rightarrow Cu + CuI_2$, the reaction is **[RPMT 1997]**

(a) Redox (b) Neutralisation
(c) Oxidation (d) Reduction

Ordinary Thinking

Objective Questions

Oxidation, Reduction

1. H_2O_2 reduces MnO_4^- ion to **[KCET (Med.) 2000]**

(a) Mn^+ (b) Mn^{2+}

(c) Mn^{3+} (d) Mn^-

2. When a sulphur atom becomes a sulphide ion **[AMU 1999]**

(a) There is no change in the composition of atom

(b) It gains two electrons

(c) The mass number changes

(d) None of these

3. The ultimate products of oxidation of most of hydrogen and carbon in food stuffs are **[DCE 2001]**

(a) H_2O alone (b) CO_2 alone

(c) H_2O and CO_2 (d) None of these

4. When P reacts with caustic soda, the products are PH_3 and NaH_2PO_2. This reaction is an example of **[IIT 1980; Kurukshetra CEE 1993; CPMT 1997]**

(a) Oxidation

(b) Reduction

(c) Oxidation and reduction (Redox)

(d) Neutralization

5. Which one of the following does not get oxidised by bromine water **[MP PET/PMT 1988]**

(a) Fe^{+2} to Fe^{+3} (b) Cu^+ to Cu^{+2}

(c) Mn^{+2} to MnO_4^- (d) Sn^{+2} to Sn^{+4}

6. In the reaction $H_2S + NO_2 \rightarrow H_2O + NO + S$. H_2S is

(a) Oxidised (b) Reduced

(c) Precipitated (d) None of these

7. The conversion of PbO_2 to $Pb(NO_3)_2$ is

(a) Oxidation

(b) Reduction

(c) Neither oxidation nor reduction

(d) Both oxidation and reduction

8. In the course of a chemical reaction an oxidant **[MP PMT 1986]**

(a) Loses electrons

(b) Gains electrons

(c) Both loses and gains electron

(d) Electron change takes place

9. $2CuI \rightarrow Cu + CuI_2$, the reaction is **[RPMT 1997]**

(a) Redox (b) Neutralisation

(c) Oxidation (d) Reduction

10. H_2S reacts with halogens, the halogens **[JIPMER 2000]**

(a) Form sulphur halides (b) Are oxidised

(c) Are reduced (d) None of these

11. H_2O_2 reduces $K_4Fe(CN)_6$ **[MP PMT 1985]**

(a) In neutral solution (b) In acidic solution

(c) In non-polar solvent (d) In alkaline solution

12. Max. number of moles of electrons taken up by one mole of NO_3^- when it is reduced to **[DPMT 2002]**

(a) NH_3 (b) NH_2OH

(c) NO (d) NO_2

13. In the reaction $3Mg + N_2 \rightarrow Mg_3N_2$ **[MP PMT 1999]**

(a) Magnesium is reduced (b) Magnesium is oxidized

(c) Nitrogen is oxidized (d) None of these

14. When sodium metal is dissolved in liquid ammonia, blue colour solution is formed. The blue colour is due to **[NCERT 1981]**

(a) Solvated Na^+ ions (b) Solvated electrons

(c) Solvated NH_2^- ions (d) Solvated protons

15. Following reaction describes the rusting of iron

$4Fe + 3O_2 \rightarrow 4Fe^{3+} + 6O^{2-}$

Which one of the following statement is incorrect **[NCERT 1981; MNR 1991; AIIMS 1998]**

(a) This is an example of a redox reaction

(b) Metallic iron is reduced to Fe^{3+}

(c) Fe^{3+} is an oxidising agent

(d) Metallic iron is a reducing agent

16. $SnCl_2$ gives a precipitate with a solution of $HgCl_2$. In this process $HgCl_2$ is **[CPMT 1983]**

(a) Reduced

(b) Oxidised

(c) Converted into a complex compound containing both Sn and Hg

(d) Converted into a chloro complex of Hg

17. Oxidation involves **[NCERT 1971, 81; CPMT 1980, 82, 83; MP PMT 1983]**

(a) Loss of electrons

(b) Gain of electrons

(c) Increase in the valency of negative part

(d) Decrease in the valency of positive part

18. Incorrect statement regarding rusting is **[MP PET 2000]**

(a) Metallic iron is oxidised to Fe^{3+} ions

(b) Metallic iron is reduced to Fe^{2-} ions

(c) Oxygen gas is reduced to oxide ion

(d) Yellowish – brown product is formed

19. When copper turnings are added to silver nitrate solution, a blue coloured solution is formed after some time. It is because, copper **[CPMT 1974, 79; DPMT 200**

(a) Displaces silver from the solution

(b) Forms a blue coloured complex with $AgNO_3$

(c) Is oxidised to Cu^{2+}

(d) Is reduced to Cu^{2+}

20. Solution of sodium metal in liquid ammonia is strongly reducing due to the presence of the following in the solution **[NCERT 1977; KCET (Med.) 2000]**

(a) Sodium atoms (b) Solvated electrons

(c) Sodium hydride (d) Sodium amide

21. When Sn^{2+} changes to Sn^{4+} in a reaction **[CPMT 1981]**

(a) It loses two electrons (b) It gains two electrons

(c) It loses two protons (d) It gains two protons

22. Oxidation of thiosulphate $(S_2O_3^{2-})$ ion by iodine gives **[NCERT 1976]**

(a) SO_3^{2-} (b) SO_4^{2-}

(c) $S_4O_6^{2-}$ (d) $S_2O_6^{2-}$

23. $Zn^{2+}(aq) + 2e \rightarrow Zn(s)$. This is **[CPMT 1985]**

(a) Oxidation (b) Reduction

(c) Redox reaction (d) None of these

24. One gas bleaches the colour of flowers by reduction while the other by oxidation **[EAMCET 1980]**

(a) CO and Cl_2 (b) SO_2 and Cl_2

(c) H_2S and Br_2 (d) NH_3 and SO_2

25. Reduction involves **NCERT 1972]**

(a) Loss of electrons

(b) Gain of electrons

(c) Increase in the valency of positive part

(d) Decrease in the valency of negative part

26. In a reaction between zinc and iodine, in which zinc iodide is formed, what is being oxidised **[NCERT 1975]**

(a) Zinc ions (b) Iodide ions

(c) Zinc atom (d) Iodine

27. Which one of the following reactions does not involve either oxidation or reduction **[EAMCET 1982]**

(a) $VO_2^+ \rightarrow V_2O_3$ (b) $Na \rightarrow Na^+$

(c) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$ (d) $Zn^{2+} \rightarrow Zn$

28. In the following reaction,

$3Br_2 + 6CO_3^{2-} + 3H_2O = 5Br^- + BrO_3^- + 6HCO_3$ **[MP PMT 1994, 95]**

(a) Bromine is oxidised and carbonate is reduced

(b) Bromine is reduced and water is oxidised

(c) Bromine is neither reduced nor oxidised

(d) Bromine is both reduced and oxidised

29. In the following reaction,

$4P + 3KOH + 3H_2O \rightarrow 3KH_2PO_2 + PH_3$ **[Pb. PMT 2002]**

(a) P is oxidized as well as reduced

(b) P is reduced only

(c) P is oxidised only

(d) None of these

30. In the following reaction

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

Which element is reduced **[CPMT 1976]**

(a) Cr (b) H

(c) O (d) I

31. The conversion of sugar $C_{12}H_{22}O_{11} \rightarrow CO_2$ is

(a) Oxidation

(b) Reduction

(c) Neither oxidation nor reduction

(d) Both oxidation and reduction

32. Which halide is not oxidised by MnO_2 **[MNR 1985; JIPMER 2000]**

(a) F (b) Cl

(c) Br (d) I

33. When Fe^{2+} changes to Fe^{3+} in a reaction

(a) It loses an electron (b) It gains an electron

(c) It loses a proton (d) It gains a proton

34. In acid solution, the reaction $MnO_4^- \rightarrow Mn^{2+}$ involves **[MP PMT 1989]**

(a) Oxidation by 3 electrons

(b) Reduction by 3 electrons

(c) Oxidation by 5 electrons

(d) Reduction by 5 electrons

35. When iron or zinc is added to $CuSO_4$ solution, copper is precipitated. It is due to **[CPMT 1974, 79]**

(a) Oxidation of Cu^{+2} (b) Reduction of Cu^{+2}

(c) Hydrolysis of $CuSO_4$ (d) Ionization of $CuSO_4$

36. In the reaction, $4Fe + 3O_2 \rightarrow 4Fe^{3+} + 6O^{2-}$ which of the following statement is incorrect **[UPSEAT 2001, 02]**

(a) A Redox reaction

(b) Metallic iron is a reducing agent

(c) Fe^{3+} is an oxidising agent

(d) Metallic iron is reduced to Fe^{3+}

37. Which of the following is redox reaction **[CBSE PMT 1997]**

(a) H_2SO_4 with $NaOH$

(b) In atmosphere, O_3 from O_2 by lightning

(c) Evaporation of H_2O

(d) Nitrogen oxides form nitrogen and oxygen by lightning

Oxidizing and Reducing agent

1. Equation $H_2S + H_2O_2 \rightarrow S + 2H_2O$ represents **[UPSEAT 2001]**

(a) Acidic nature of H_2O_2

(b) Basic nature of H_2O_2

(c) Oxidising nature of H_2O_2

(d) Reducing nature of H_2O_2

2. In the reaction

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

the reductant is **[EAMCET 1991]**

(a) $C_2O_4^{2-}$ (b) MnO_4^-

(c) Mn^{2+} (d) H^+

3. A reducing agent is a substance which can **[CPMT 1971, 74, 76, 78, 80; NCERT 1976]**

(a) Accept electron (b) Donate electrons

(c) Accept protons (d) Donate protons

4. Which of the following is the most powerful oxidizing agent **[MNR 1990; CPMT 2003]**

(a) F_2 (b) Cl_2

(c) Br_2 (d) I_2

5. Of the four oxyacids of chlorine the strongest oxidising agent in dilute aqueous solution is **[MP PET 2000]**

(a) $HClO_4$ (b) $HClO_3$

(c) $HClO_2$ (d) $HOCl$

6. Identify the correct statement about H_2O_2 **[AIIMS 1996]**

(a) It acts as reducing agent only

(b) It acts as both oxidising and reducing agent

(c) It is neither an oxidiser nor reducer

(d) It acts as oxidising agent only

7. Several blocks of magnesium are fixed to the bottom of a ship to **[AIEEE 2003]**

(a) Keep away the sharks

(b) Make the ship lighter

(c) Prevent action of water and salt

(d) Prevent puncturing by under-sea rocks

8. Which of the following behaves as both oxidising and reducing agents **[AFMC 1995]**

(a) H_2SO_4 (b) SO_2

(c) H_2S (d) HNO_3

9. The reaction $H_2S + H_2O_2 \rightarrow 2H_2O + S$ shows **[JIPMER 2001]**

(a) Oxidizing action of H_2O_2

(b) Reducing action of H_2O_2

(c) Alkaline nature of H_2O_2

(d) Acidic nature of H_2O_2

10. Which of the following is not a reducing agent **[EAMCET 1987]**

(a) $NaNO_2$ (b) $NaNO_3$

(c) HI (d) $SnCl_2$

11. Which of the following cannot work as oxidising agent **[CPMT 1996]**

(a) O_2 (b) $KMnO_4$

(c) I_2 (d) None of these

12. H_2O_2 is used as **[CPMT 1994]**

(a) An oxidant only

(b) A reductant only

(c) An acid only

(d) An oxidant, a reductant and an acid

13. In $C + H_2O \rightarrow CO + H_2$, H_2O acts as **[AFMC 1988]**

(a) Oxidising agent (b) Reducing agent

(c) (a) and (b) both (d) None of these

14. Strongest reducing agent is **[CPMT 1977; BHU 1984, 96; MP PET 1990; AMU 1999]**

(a) F^- (b) Cl^-

(c) Br^- (d) I^-

15. A solution of sulphur dioxide in water reacts with H_2S precipitating sulphur. Here sulphur dioxide acts as **[NCERT 1980]**

(a) As oxidising agent (b) A reducing agent

(c) An acid (d) A catalyst

16. Which of these substances is a good reducing agent **[NCERT 1979; CPMT 1988]**

(a) $NaOCl$ (b) HI

(c) $FeCl_3$ (d) KBr

17. The strongest reducing agent is **[MNR 1982]**

(a) HNO_2 (b) H_2S

(c) H_2SO_3 (d) $SnCl_2$

18. Which one is an oxidising agent **[DPMT 1996]**

(a) $FeSO_4$

(b) HNO_3

(c) $FeSO_4.(NH_4)_2SO_4.6H_2O$

(d) H_2SO_4

19. In which of the following reactions H_2O_2 is a reducing agent **[CPMT 1981; NCERT 1981; BHU 1999]**

(a) $2FeCl_2 + 2HCl + H_2O_2 \rightarrow 2FeCl_3 + 2H_2O$

(b) $Cl_2 + H_2O_2 \rightarrow 2HCl + O_2$

(c) $2HI + H_2O_2 \rightarrow 2H_2O + I_2$

(d) $H_2SO_3 + H_2O_2 \rightarrow H_2SO_4 + H_2O$

20. When $NaCl$ is dissolved in water the sodium ion becomes **[NCERT 1976]**

(a) Oxidised (b) Reduced

(c) Hydrolysed (d) Hydrated

21. Strongest reducing agent is **[MNR 1984, 89]**

(a) K (b) Mg

(c) Al (d) Br

(e) Na

22. Which substance is serving as a reducing agent in the following reaction

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

[CBSE PMT 1994; AFMC 2000; DPMT 2001]

(a) H_2O (b) Ni

(c) H^+ (d) $Cr_2O_7^{2-}$

23. Which of the following acid possesses oxidising, reducing and complex forming properties **[MNR 1985]**

(a) HNO_3 (b) H_2SO_4

(c) HCl (d) HNO_2

24. Which one is oxidising substance **[CPMT 1997]**

(a) $C_2H_2O_2$ (b) CO

(c) H_2S (d) CO_2

25. The compound that can work both as oxidising and reducing agent is **[CPMT 1986; MP PET 2000]**

(a) $KMnO_4$ (b) H_2O_2

(c) BaO_2 (d) $K_2Cr_2O_7$

26. Which one is oxidising agent in the reaction below

$2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$ **[CPMT 1997]**

(a) H^+ (b) $Cr_2O_4^-$

(c) Cr^{++} (d) None of these

27. Which is the best description of the behaviour of bromine in the reaction given below

$H_2O + Br_2 \rightarrow HOBr + HBr$ **[CBSE PMT 2004]**

(a) Oxidised only

(b) Reduced only

(c) Proton acceptor only

(d) Both oxidised and reduced

28. What is the oxidising agent in chlorine water

[JEE Orissa 2004]

(a) HCl (b) $HClO_2$

(c) $HOCl$ (d) None of these

29. In the reaction

$Ag_2O + H_2O_2 \rightarrow 2Ag + H_2O + O_2$, the H_2O_2 acts as

[BHU 2004]

(a) Reducing agent (b) Oxidising agent

(c) Bleaching agent (d) None of the above

30. In the reaction

$HAsO_2 + Sn^{2+} \rightarrow As + Sn^{4+} + H_2O$ oxidising agent is

[BVP 2004]

(a) Sn^{2+} (b) Sn^{4+}

(c) As (d) $HAsO_2$

31. Which of the following substances acts as an oxidising as well as a reducing agent **[UPSEAT 2004; DCE 20**

(a) Na_2O (b) $SnCl_2$

(c) Na_2O_2 (d) $NaNO_2$

32. In the reaction

$P + NaOH \rightarrow PH_3 + NaH_2PO_2$ **[MP PET 2004]**

(a) P is oxidised only

(b) P is reduced only

(c) P is oxidized as well as reduced

(d) Na is reduced

Oxidation number and Oxidation state

1. The oxidation number of C in CO_2 is **[MP PET 2001]**

(a) – 2 (b) + 2

(c) – 4 (d) + 4

2. The oxidation number of As is **[RPMT 1997]**

(a) + 2 and + 3 (b) + 3 and + 5

(c) + 3 and + 4 (d) None of these

3. The oxidation number of Ba in barium peroxide is

[Pb. PMT 2002]

(a) + 6 (b) + 2

(c) 1 (d) + 4

4. HNO_2 acts both as reductant and oxidant, while HNO_3 acts only as oxidant. It is due to their **[AIIMS 2000]**

(a) Solubility ability

(b) Maximum oxidation number

(c) Minimum oxidation number

(d) Minimum number of valence electrons

5. Chlorine is in +1 oxidation state in

[MP PMT 1981; NCERT 1974; CPMT 1971, 78]

(a) HCl (b) $HClO_4$

(c) ICl (d) Cl_2O

6. The valency of Cr in the complex $[Cr(H_2O)_4Cl_2]^+$

[MP PMT 2000]

(a) 1 (b) 3

(c) 5 (d) 6

7. In the conversion $Br_2 \rightarrow BrO_3^-$, the oxidation state of bromine changes from

[EAMCET 1990; AMU 1999; RPMT 2002]

(a) – 1 to – 1 (b) 0 to – 1

(c) 0 to + 5 (d) 0 to – 5

8. In the chemical reaction $Cl_2 + H_2S \rightarrow 2HCl + S$, the oxidation number of sulphur changes from **[MP PMT 1999]**

(a) 0 to 2 (b) 2 to 0

(c) – 2 to 0 (d) – 2 to – 1

9. Oxidation number of cobalt in $K[Co(CO)_4]$ is

[KCET 1996]

(a) + 1 (b) + 3

(c) – 1 (d) – 3

10. When $K_2Cr_2O_7$ is converted to K_2CrO_4, the change in the oxidation state of chromium is **[NCERT 1981]**

(a) 0 (b) 6
(c) 4 (d) 3

11. The oxidation number of chlorine in $HOCl$

(a) – 1 (b) 0
(c) + 1 (d) + 2

12. Oxidation number of S in S^{2-} is **[CPMT 1979]**

(a) – 2 (b) 0
(c) – 6 (d) + 2

13. Oxidation number of N in $(NH_4)_2SO_4$ is **[CPMT 1996]**

(a) – 1 / 3 (b) – 1
(c) + 1 (d) – 3

14. In which compound, oxidation state of nitrogen is 1

[MP PMT 1989]

(a) NO (b) N_2O
(c) NH_2OH (d) N_2H_4

15. Oxidation number of nickel in $Ni(CO)_4$

[AIIMS 1984; MNR 1985; CPMT 1997; MP PET/PMT 1998; AMU 2000; 01]

(a) 0 (b) + 4
(c) – 4 (d) + 2

16. The oxidation number of sulphur in H_2SO_4 is

[CPMT 1979Pb. CET 2002]

(a) – 2 (b) + 2
(c) + 4 (d) + 6

17. Oxidation state of chlorine in perchloric acid is

[EAMCET 1989]

(a) – 1 (b) 0
(c) – 7 (d) + 7

18. Oxidation number of N in HNO_3 is

[BHU 1997]

(a) – 3.5 (b) + 3.5
(c) – 3, +5 (d) + 5

19. The oxidation number of Mn in MnO_4^{-1} is

(a) + 7 (b) – 5
(c) + 6 (d) + 5

20. Sn^{++} loses two electrons in a reaction. What will be the oxidation number of tin after the reaction

(a) + 2 (b) Zero
(c) + 4 (d) – 2

21. The oxidation state of Mn in K_2MnO_4

[CPMT 1982, 83, 84; DPMT 1982; NCERT 1973; AMU 2000]

(a) + 2 (b) + 7
(c) – 2 (d) + 6

22. Oxidation number of oxygen in O_2 molecule is

[CPMT 1984]

(a) + 1 (b) 0
(c) + 2 (d) – 2

23. Maximum oxidation state of Cr is **[RPMT 2002]**

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

24. In which of the following compound transition metal has zero oxidation state **[CBSE PMT 1999; BHU 2000]**

(a) CrO_5 (b) $NH_2.NH_2$
(c) $NOClO_4$ (d) $[Fe(CO)_5]$

25. Carbon is in the lowest oxidation state in

[NCERT 1979; MH CET 1999]

(a) CH_4 (b) CCl_4
(c) CF_4 (d) CO_2

26. Oxidation number of carbon in $H_2C_2O_4$ is

[CPMT 1982]

(a) + 4 (b) + 3
(c) + 2 (d) – 2

27. The oxidation number of Pt in $[Pt(C_2H_4)Cl_3]^-$ is

[MNR 1993]

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

28. The oxidation number of carbon in CH_2Cl_2 is

[CPMT 1976; Pb. PET 1999; AFMC 2004]

(a) 0 (b) + 2
(c) -2 (d) + 4

29. The oxidation states of phosphorus vary from

[CPMT 1976]

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

30. The process in which oxidation number increases is known as

[CPMT 1976]

(a) Oxidation (b) Reduction
(c) Auto-oxidation (d) None of the above

31. The oxidation number of S in $H_2S_2O_8$ is **[MP PET 2002]**

(a) + 2 (b) + 4
(c) + 6 (d) + 7

32. The oxidation state of nitrogen in N_3H is

[NCERT 1977, 81]

(a) $+\frac{1}{3}$ (b) + 3
(c) -1 (d) $-\frac{1}{3}$

33. Which of the following statements is correct **[AFMC 1997]**

(a) Hydrogen has oxidation number -1 and $+1$
(b) Hydrogen has same electronegativity as halogens
(c) Hydrogen will not be liberated at anode
(d) Hydrogen has same ionization potential as alkali metals

34. The oxidation state of Cr in $[Cr(NH_3)_4Cl_2]^+$ is

[AIEEE 2005]

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

35. Sulphur has highest oxidation state in [EAMCET 1991]
(a) SO_2 (b) H_2SO_4
(c) $Na_2S_2O_3$ (d) $Na_2S_4O_6$

36. The oxidation number of Fe and S in iron pyrites are [RPMT 1997]
(a) 4, – 2 (b) 2, – 1
(c) 3, – 1.5 (d) 3, – 1

37. The oxidation number of nitrogen in NO_3^- is [CPMT 1982]
(a) – 1 (b) + 2
(c) + 3 (d) + 5

38. Oxidation state of elemental carbon is [MNR 1983]
(a) 0 (b) 1
(c) 2 (d) 3

39. The sum of the oxidation numbers of all the carbons in C_6H_5CHO is [EAMCET 1986]
(a) + 2 (b) 0
(c) + 4 (d) – 4

40. Which one of the following has the highest oxidation number of iodine [CPMT 1982]
(a) KI_3 (b) KI
(c) IF_5 (d) KIO_4

41. The oxidation number of N in $N_2H_5^+$ [Pb. PMT 2001]
(a) – 3 (b) (– 2)
(c) – 1 (d) + 2

42. In which of the following compounds the oxidation number of carbon is maximum
(a) $HCHO$ (b) $CHCl_3$
(c) CH_3OH (d) $C_{12}H_{22}O_{11}$

43. The oxidation state of chlorine in $KClO_4$ is [CPMT 1985]
(a) – 1 (b) + 1
(c) + 7 (d) – 7

44. The oxidation state of I in $H_4IO_6^-$ is [CBSE PMT 1994]
(a) + 7 (b) + 5
(c) + 1 (d) – 1

45. An element which never has a positive oxidation number in any of its compounds [AIIMS 1981]
(a) Boron (b) Oxygen
(c) Chlorine (d) Fluorine

46. In an oxidation process, oxidation number [CPMT 1976]
(a) Decreases
(b) Increases
(c) Does not change
(d) First increases then decreases

47. If HNO_3 changes into N_2O, the oxidation number is changed by [BHU 1997; AFMC 2001]
(a) + 2 (b) – 1
(c) 0 (d) + 4

48. The characteristic oxidation number of atoms in free metals is [NCERT 1975]
(a) Minus one (b) Any number
(c) One (d) Zero

49. In which one of the following changes there are transfer of five electrons [NCERT 1982]
(a) $MnO_4^- \rightarrow Mn^{2+}$ (b) $CrO_4^2 \rightarrow Cr^{3+}$
(c) $MnO_4^{2-} \rightarrow MnO_2$ (d) $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

50. Oxidation number of C in $C_6H_{12}O_6$ is [KCET 1992]
(a) + 6 (b) – 6
(c) 0 (d) + 4

51. In which of the following compounds iron has lowest oxidation state [MNR 1984]
(a) $FeSO_4.(NH_4)_2SO_4.6H_2O$
(b) $K_4Fe(CN)_6$
(c) $Fe(CO)_5$
(d) Fe_2O
(e) K_2FeO_4

52. The oxidation number of hydrogen in MH_2 is [CPMT 1976]
(a) + 1 (b) – 1
(c) + 2 (d) – 2

53. Oxidation number of iodine varies from [CPMT 1982]
(a) – 1 to +1 (b) – 1 to +7
(c) +3 to +5 (d) – 1 to +5

54. When SO_2 is passed through acidic solution of potassium dichromate, then chromium sulphate is formed. Change in valency of chromium is [CPMT 1979]
(a) +4 to +2 (b) +5 to +3
(c) +6 to +3 (d) +7 to +2

55. The oxidation states of the most electronegative element in the products of the reaction of BaO_2 with dilute H_2SO_4 are [IIT 1991; CBSE PMT 1992; BHU 2000]
(a) 0 and – 1 (b) – 1 and – 2
(c) – 2 and 0 (d) – 2 and + 1

56. The highest oxidation state of Mn is shown by [MNR 1983; RPMT 1999]
(a) K_2MnO_4 (b) $KMnO_4$
(c) MnO_2 (d) Mn_2O_2
(e) MnO

57. The oxidation number of carbon in CH_2O is [IIT 1982; EAMCET 1985; MNR 1990; UPSEAT 2001 CPMT 1997, 2004]
(a) – 2 (b) + 2
(c) 0 (d) + 4

58. Oxidation state of oxygen in hydrogen peroxide is [DPMT 1984; 91; CPMT 1988; MNR 1994; UPSEAT 2001; RPMT 2002; JEE Orissa 2004]
(a) – 1 (b) + 1
(c) 0 (d) – 2

59. The oxidation number of Cr in $K_2Cr_2O_7$ is

[CPMT 1981, 85, 90, 93, 99; KCET 1992; BHU 1988, 98; AFMC 1991, 99; EAMCET 1986; MP PMT 1996, 99, 2002; MP PET/PMT 1998; Bihar CEE 1995; RPET 2000]

(a) +6 (b) – 7
(c) +2 (d) – 2

60. In which of the following compounds transition metal is in oxidation state zero **[NCERT 1982]**

(a) $[Co(NH_3)_6]Cl_2$ (b) $[Fe(H_2O)_6SO_4]$
(c) $[Ni(CO)_4]$ (d) $[Fe(H_2O)_3](OH)_2$

61. Oxidation number of osmium (Os) in OsO_4 is

[AIIMS 1999]

(a) + 4 (b) + 6
(c) + 7 (d) + 8

62. The atomic number of an element which shows the oxidation state of + 3 is **[CPMT 1989, 94]**

(a) 13 (b) 32
(c) 33 (d) 17

63. The oxidation number of iron in the compound $K_4[Fe(CN)_6]$ is

[NCERT 1976; MNR 1986; AIIMS 2000]

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

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

[EAMCET 1987; IIT 1987; MP PMT 1994; AIIMS 1997; DCE 2000]

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

65. Oxidation state of oxygen in F_2O is

[BHU 1982; UPSEAT 2001; MH CET 2002]

(a) + 1 (b) + 2
(c) –1 (d) –2

66. Phosphorus has the oxidation state of +3 in

[NCERT 1982; RPMT 1999]

(a) Orthophosphoric acid (b) Phosphorus acid
(c) Metaphosphoric acid (d) Pyrophosphoric acid

67. Oxidation number of P in $Mg_2P_2O_7$ is

[CPMT 1989; MP PMT 1995]

(a) + 3 (b) + 2
(c) + 5 (d) – 3

68. The oxidation state of nitrogen is highest in

[MP PMT 2001; BHU 2002]

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

69. Oxidation number of P in KH_2PO_2 is

[CPMT 1987; MH CET 1999]

(a) + 1 (b) + 3
(c) + 5 (d) – 4

70. The most common oxidation state of an element is –2. The number of electrons present in its outermost shell is

[BHU 1983; NCERT 1974; CPMT 1977]

(a) 4 (b) 2
(c) 6 (d) 8

71. Sulphur has lowest oxidation number in

[EAMCET 1993]

(a) H_2SO_3 (b) SO_2
(c) H_2SO_4 (d) H_2S

72. The oxidation number and covalency of sulphur in the sulphur molecule (S_8) are respectively **[NCERT 1977]**

(a) 0 and 2 (b) 6 and 8
(c) 0 and 8 (d) 6 and 2

73. In ferrous ammonium sulphate oxidation number of Fe is

[CPMT 1988]

(a) + 3 (b) + 2
(c) + 1 (d) – 2

74. The oxidation number of nitrogen in NH_2OH is

[NCERT 1981]

(a) + 1 (b) – 1
(c) – 3 (d) – 2

75. The oxidation number of phosphorus in $Ba(H_2PO_2)_2$ is

[Kurukshetra CEE 1998; DCE 2004]

(a) – 1 (b) + 1
(c) + 2 (d) + 3

76. A compound is in its low oxidation state. Then its will be

[DCE 2001]

(a) Highly acidic
(b) Highly basic
(c) Highest oxidising property
(d) Half acidic, half basic

77. The oxidation number and the electronic configuration of sulphur in H_2SO_4 is **[KCET 2002]**

(a) + 4; $1s^2 2s^2 2p^6 3s^2$
(b) + 2; $1s^2 2s^2 2p^6 3s^2 3p^2$
(c) + 3; $1s^2 2s^2 2p^6 3s^2 3p^1$
(d) + 6; $1s^2 2s^2 2p^6$

78. The oxidation number of Mn in $KMnO_4$ is

[CPMT 1982, 83; EAMCET 1992, 93; RPET 1999]

(a) + 7 (b) – 7
(c) + 1 (d) – 1

79. Oxidation number of As atoms in H_3AsO_4 is

[DPMT 2001]

(a) – 3 (b) + 4
(c) + 6 (d) + 5

80. In XeO_3 and XeF_6 the oxidation state of Xe is

[MP PET 2003]

(a) + 4 (b) + 6
(c) + 1 (d) + 3

81. Oxidation number of carbon in CH_3-Cl is

[MP PET 2000]

(a) – 3 (b) – 2
(c) – 1 (d) 0

82. The oxidation state of Cr in $Cr_2O_7^{2-}$ is [BHU 2000; CPMT 2000]

(a) 4 (b) – 6
(c) 6 (d) – 2

83. Oxidation state of 'S' in H_2SO_3 [RPET 2003]

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

84. Oxidation numbers of two Cl atoms in bleaching powder, $CaOCl_2$ are

(a) – 1, – 1 (b) + 1, – 1
(c) + 1, + 1 (d) 0, – 1

85. Select the compound in which chlorine is assigned the oxidation number +5 [NCERT 1984, 94]

(a) $HClO_4$ (b) $HClO_2$
(c) $HClO_3$ (d) HCl

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

[MNR 1987; MP PET 2000; CBSE PMT 2000; UPSEAT 2000, 02; BHU 2003; AMU 2002]

(a) 7 to 4 (b) 6 to 4
(c) 7 to 2 (d) 4 to 2

87. Oxygen has oxidation states of +2 in the [NCERT 1973; DPMT 1983; MP PET 2000]

(a) H_2O_2 (b) CO_2
(c) H_2O (d) OF_2

88. The element exhibiting most stable $+2$ oxidation state among the following is [IIT 1995]

(a) Ag (b) Fe
(c) Sn (d) Pb

89. Oxidation number of sulphur in $S_2O_3^{2-}$ is[CPMT 1979]

(a) – 2 (b) + 2
(c) + 6 (d) 0

90. Carbon has zero oxidation number in [Kurukshetra CEE 2002]

(a) CO (b) CH_4
(c) CH_2Cl_2 (d) CH_3Cl

91. Oxidation state of oxygen atom in potassium superoxide is [MNR 1988; NCERT 1980]

(a) 0 (b) – 1
(c) $-\frac{1}{2}$ (d) – 2

92. Oxidation number of S in S_2Cl_2 is

(a) + 1 (b) – 1
(c) + 6 (d) 0

93. What is the oxidation number of sulphur in $Na_2S_4O_6$ [AIIMS 1998; DCE 1999]

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

94. When $CuSO_4$ reacts with KI, the oxidation number of Cu changes by [BHU 1997]

(a) 0 (b) – 1
(c) 1 (d) 2

95. The oxidation number of N in NH_4Cl is

(a) + 5 (b) + 3
(c) – 5 (d) – 3

96. In which reaction there is a change in valency [NCERT 1971; CPMT 1971]

(a) $2NO_2 \rightarrow N_2O_4$
(b) $2NO_2 + H_2O \rightarrow HNO_2 + HNO_3$
(c) $NH_4OH \rightarrow NH_4^+ + OH^-$
(d) $CaCO_3 \rightarrow CaO + CO_2$

97. Oxidation state of Fe in Fe_3O_4 is [CBSE PMT 1999; AIIMS 2002]

(a) $\frac{3}{2}$ (b) $\frac{4}{5}$
(c) $\frac{5}{4}$ (d) $\frac{8}{3}$

98. Nitrogen show different oxidation states in the range [Kerala (Med.) 2003]

(a) 0 to +5 (b) – 3 to + 5
(c) – 5 to + 3 (d) – 3 to + 3

99. Oxidation number of Mn in K_2MnO_4 and $MnSO_4$ are respectively [CPMT 1997]

(a) + 7, + 2 (b) + 6, + 2
(c) + 5, + 2 (d) + 2, + 6

100. Identify the element which can have highest oxidation numbers [AIIMS 1996]

(a) N (b) O
(c) Cl (d) C

101. What is the oxidation number of Co in $[Co(NH_3)_4ClNO_2]$ [BHU 1999]

(a) + 2 (b) + 3
(c) + 4 (d) + 5

102. The oxidation number of nickel in $K_4[Ni(CN)_4]$ is [JIPMER 1999]

(a) – 2 (b) – 1
(c) + 2 (d) 0

103. The oxidation number of fluorine in F_2O is [CPMT 1982; BHU 1982; EAMCET 1986]

(a) – 1 (b) + 1
(c) + 2 (d) – 2

104. Oxidation number of Fe in $K_3[Fe(CN)_6]$ is [AMU 1988]

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

105. Oxidation number of N in NH_3 is **[CPMT 1979; Pb CET 2004]**
(a) – 3 (b) + 3
(c) 0 (d) + 5

106. What is the net charge on ferrous ion **[AFMC 2004]**
(a) + 2 (b) + 3
(c) + 4 (d) + 5

107. Which of the following elements never show positive oxidation number **[CPMT 2004]**
(a) O (b) Fe
(c) Ga (d) F

108. The oxidation state shown by silicon when it combines with strongly electropositive metals is **[MH CET 2004]**
(a) – 2 (b) – 4
(c) + 4 (d) – 2

109. The oxidation number of sulphur in H_2S is **[Pb. CET 2002]**
(a) – 2 (b) + 3
(c) + 2 (d) – 3

110. Oxidation number of nitrogen in $NaNO_2$ is **[Pb. CET 2000]**
(a) + 2 (b) + 3
(c) + 4 (d) – 3

111. Oxidation number of S in SO_4^{2-} **[BCECE 2005]**
(a) + 6 (b) + 3
(c) + 2 (d) – 2

112. The oxidation state of chromium in the final product formed by the reaction between KI and acidified potassium dichromate solution is **[AIEEE 2005]**
(a) +4 (b) +6
(c) +2 (d) +3

113. The oxidation state of I in IPO_4 is **[Orissa JEE 2005]**
(a) +1 (b) +3
(c) +5 (d) +7

Redox reaction and Method for balancing Redox reaction

1. The value of x in the partial redox equation $MnO_4^- + 8H^+ + xe \rightleftharpoons Mn^{2+} + 4H_2O$ is
(a) 5 (b) 3
(c) 1 (d) 0

2. $C_2H_6(g) + nO_2 \rightarrow CO_2(g) + H_2O(l)$
In this equation, the ratio of the coefficients of CO_2 and H_2O is **[KCET 1992]**
(a) 1 : 1 (b) 2 : 3
(c) 3 : 2 (d) 1 : 3

3. The number of electrons involved in the reduction of $Cr_2O_7^{2-}$ in acidic solution to Cr^{3+} is **[EAMCET 1983]**
(a) 0 (b) 2
(c) 3 (d) 5

4. When $KMnO_4$ acts as an oxidising agent and ultimately forms $[MnO_4]^{-2}$, MnO_2, Mn_2O_3, Mn^{+2} then the number of electrons transferred in each case respectively is **[AIEEE 2002]**
(a) 4, 3, 1, 5 (b) 1, 5, 3, 7
(c) 1, 3, 4, 5 (d) 3, 5, 7, 1

5. Starch paper is used to test for the presence of **[NCERT 1979]**
(a) Iodine (b) Oxidising agent
(c) Iodide ion (d) Reducing agent

6. How many moles of $K_2Cr_2O_7$ can be reduced by 1 mole of Sn^{2+} **[MP PMT 2003]**
(a) 1/3 (b) 1/6
(c) 2/3 (d) 1

7. $2MnO_4^- + 5H_2O_2 + 6H^+ \rightarrow 2\,Z + 5O_2 + 8H_2O$. In this reaction Z is **[RPMT 2002]**
(a) Mn^{+2} (b) Mn^{+4}
(c) MnO_2 (d) Mn

8. What is 'A' in the following reaction
$2Fe^{3+}{}_{(aq)} + Sn^{2+}{}_{(aq)} \rightarrow 2Fe^{2+}{}_{(aq)} + A$ **[MP PET 2003]**
(a) $Sn^{3+}{}_{(aq)}$ (b) $Sn^{4+}{}_{(aq)}$
(c) $Sn^{2+}{}_{(aq)}$ (d) Sn

9. 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 reaction are **[IIT 1988, 92; BHU 1995; CPMT 1997; RPMT 1999; DCE 2000; MP PET 2003]**

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

10. Which of the following is a redox reaction **[AIEEE 2002]**
(a) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$
(b) $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$
(c) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2NH_4OH$
(d) $Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$

11. Which of the following reaction is a redox reaction **[MP PMT 2003]**
(a) $P_2O_5 + 2H_2O \rightarrow H_4P_2O_7$
(b) $2AgNO_3 + BaCl_2 \rightarrow 2AgCl + Ba(NO_3)_2$
(c) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$
(d) $Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$

12. Which of the following reactions involves oxidation-reduction

[NCERT 1972; AFMC 2000; Pb. CET 2004; CPMT 2004]

(a) $NaBr + HCl \rightarrow NaCl + HBr$

(b) $HBr + AgNO_3 \rightarrow AgBr + HNO_3$

(c) $H_2 + Br_2 \rightarrow 2HBr$

(d) $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

13. Which of the following is the strongest oxidising agent **[Pb. CET 2000]**

(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) $Cr_2O_7^{2-} / Cr^{3+}, E^o = +1.33$

14. In the balanced chemical reaction,

$IO_3^- + a\, I^- + b\, H^+ \rightarrow c\, H_2O + d\, I_2$

a, b, c and *d* respectively correspond to **[AIIMS 2005]**

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

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

15. The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is: **[CBSE PMT 2005]**

(a) One fifth (b) five

(c) One (d) Two

Auto oxidation and Disproportionation

1. In the equation $H_2S + 2HNO_3 \rightarrow 2H_2O + 2NO_2 + S$

The equivalent weight of hydrogen sulphide is **[BVP 2003]**

(a) 16 (b) 68

(c) 34 (d) 17

2. If 1.2 *g* of metal displace 1.12 litre hydrogen at normal temperature and pressure ,equivalent weight of metal would be **[DPMT 2001]**

(a) 24 (b) 12

(c) 1.2 ÷11.2 (d) 1.2 × 11.2

3. Which one of the following nitrates will leave behind a metal on strong heating **[AIEEE 2003]**

(a) Ferric nitrate (b) Copper nitrate

(c) Manganese nitrate (d) Silver nitrate

4. To prevent rancidification of food material, which of the following is added **[CPMT 1996]**

(a) Reducing agent (b) Anti-oxidant

(c) Oxidising agent (d) None of these

5. Prevention of corrosion of iron by zinc coating is called **[MP PMT 1993; CPMT 2002]**

(a) Galvanization (b) Cathodic protection

(c) Electrolysis (d) Photo-electrolysis

6. The metal used in galvanizing of iron is **[MP PET 1985, 96]**

(a) Pb (b) Zn

(c) Al (d) Sn

7. In which of the following reactions there is no change in valency **[NCERT 1974; CPMT 1978]**

(a) $4KClO_3 \rightarrow 3KClO_4 + KCl$

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

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

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

8. The equivalent weight of phosphoric acid (H_3PO_4) in the reaction $NaOH + H_3PO_4 \rightarrow NaH_2PO_4 + H_2O$ is **[AIIMS 1999]**

(a) 25 (b) 49

(c) 59 (d) 98

9. What is the equivalent mass of IO_4^- when it is converted into I_2 in acid medium **[Kerala PMT 2004]**

(a) *M*/6 (b) *M*/7

(c) *M*/5 (d) *M*/4

(e) None of these

10. For decolourization of 1 mole of $KMnO_4$, the moles of H_2O_2 required is **[AIIMS 2004]**

(a) 1/2 (b) 3/2

(c) 5/2 (d) 7/2

11. In the reaction $I_2 + 2S_2O_3^{--} \rightarrow 2I^- + S_4O_6^{--}$ equivalent weight of iodine will be equal to **[MP PET 2004]**

(a) 1/2 of molecular weight

(b) Molecular weight

(c) 1/4 of molecular weight

(d) None

12. The equivalent weight of KIO_3 in the reaction $2Cr(OH)_3 + 4OH + KIO_3 \rightarrow 2CrO_4^{2-} + 5H_2O + KI$ is **[MP PMT 2004]**

(a) Mole wt. (b) $\frac{\text{Mol.wt.}}{6}$

(c) $\frac{\text{Mol.wt.}}{2}$ (d) $\frac{\text{Mol.wt.}}{3}$

13. The product of oxidation of I^- with MnO_4^- in alkaline medium is **[IIT-JEE Screening 2004]**

(a) IO_3^- (b) I_2

(c) IO^- (d) IO_4^-

14. In alkaline medium ClO_2 oxidize H_2O_2 in O_2 and reduced itself in Cl^- then how many mole of H_2O_2 will oxidize by one mole of ClO_2 **[Kerala CET 2005]**

(a) 1.0 (b) 1.5

(c) 2.5 (d) 3.5

(e) 5.0

Critical Thinking

Objective Questions

1. In which of the following acid, which acid has oxidation reduction and complex formation properties **[UPSEAT 2001]**

(a) HNO_3 (b) H_2SO_4

(c) HCl (d) HNO_2

2. The compound which could not act both as oxidising as well as reducing agent is **[IIT Screening 1991]**

(a) SO_2 (b) MnO_2

(c) Al_2O_3 (d) CrO

3. H_2S acts only as a reducing agent while SO_2 can act both as a reducing and oxidizing agent because **[AMU 1999]**

(a) S in H_2S has – 2 oxidation state

(b) S in SO_2 has oxidation state + 4

(c) Hydrogen in H_2S more +*ve* than oxygen

(d) Oxygen is more – *ve* in SO_2

4. Of all the three common mineral acids, only sulphuric acid is found to be suitable for making the solution acidic because **[Kurukshetra CEE 2002]**

(a) It does not react with $KMnO_4$ or the reducing agent

(b) Hydrochloric acid reacts with $KMnO_4$

(c) Nitric acid is an oxidising agent which reacts with reducing agent

(d) All of the above are correct

5. For H_3PO_3 and H_3PO_4 the correct choice is **[IIT Screening 2003]**

(a) H_3PO_3 is dibasic and reducing

(b) H_3PO_3 is dibasic and non-reducing

(c) H_3PO_4 is tribasic and reducing

(d) H_3PO_3 is tribasic and non-reducing

6. Match List I with List II and select the correct answer using the codes given below the lists

List I (Compound)	List II (Oxidation state of *N*)
(A) NO_2	(1) + 5
(B) HNO	(2) – 3
(C) NH_3	(3) + 4
(D) N_2O_5	(4) + 1

Codes :

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

7. M^{+3} ion loses $3e^-$. Its oxidation number will be **[CPMT 2002]**

(a) 0 (b) + 3

(c) + 6 (d) – 3

8. In the reaction $Zn + 2H^+ + 2Cl^- \rightarrow Zn^{2+} + 2Cl^- + H_2$, the spectator ion is **[AIIMS 2001]**

(a) Cl^- (b) Zn^{2+}

(c) H^+ (d) All of these

9. The oxidation number of sulphur in $H_2S_2O_7$ and iron in $K_4Fe(CN)_6$ is respectively **[AIIMS 2000]**

(a) + 6 and + 2 (b) + 2 and + 2

(c) + 8 and + 2 (d) + 6 and + 4

10. Oxidation number of oxygen in potassium super oxide (KO_2) is **[UPSEAT 1999, 2002]**

(a) – 2 (b) – 1

(c) – 1/2 (d) – 1/4

11. One mole of N_2H_4 loses 10 mol of electrons to form a new compound Y. Assuming that all nitrogen appear in the new compound, what is the oxidation state of N_2 in Y ? (There is no change in the oxidation state of hydrogen) **[IIT 1981; Pb. PMT 1998]**

(a) + 3 (b) – 3

(c) – 1 (d) + 5

12. Amongst the following identify the species with an atom in + 6 oxidation state **[IIT Screening 2000]**

(a) MnO_4^- (b) $Cr(CN)_6^{3-}$

(c) NiF_6^{2-} (d) CrO_2Cl_2

13. In which of the following compounds, is the oxidation number of iodine is fractional **[BVP 2003]**

(a) IF_3 (b) IF_2

(c) I_3^- (d) IF_7

14. The compound $YBa_2Cu_3O_7$ which shows superconductivity has copper in oxidation state .... Assume that the rare earth element Yttrium is in its usual $+3$ oxidation state **[IIT 1994]**

(a) 3/7 (b) 7/3

(c) 3 (d) 7

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

(a) 0, +1 and – 2 (b) + 2, +1 and – 2

(c) 0, + 1 and + 2 (d) – 2, + 1 and – 2

16. Which one of the following reactions is not an example of redox reaction **[Kurukshetra CEE 1998]**

(a) $Cl_2 + 2H_2O + SO_2 \rightarrow 4H^+ + SO^{4-} + 2Cl^-$

(b) $Cu^{++} + Zn \rightarrow Zn^{++} + Cu$

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

(d) $HCl + H_2O \rightarrow H_3O^- + Cl^-$

17. For the reactions, $C + O_2 \rightarrow CO_2; \Delta H = -393\,J$

$2\,Zn + O_2 \rightarrow 2\,ZnO; \Delta H = -412\,J$ **[AIEEE 2002]**

(a) Carbon can oxidise Zn

(b) Oxidation of carbon is not feasible

(c) Oxidation of Zn is not feasible

(d) Zn can oxidise carbon

18. In the reaction $B_2H_6 + 2KOH + 2X \rightarrow 2Y + 6H_2$, X and Y are respectively **[EAMCET 2003]**

(a) H_2, H_3BO_3 (b) HCl, KBO_3

(c) H_2O, KBO_3 (d) H_2O, KBO_2

19. In a balanced equation $H_2SO_4 + x\,HI \rightarrow H_2S + y\,I_2 + z\,H_2O$, the values of *x*, *y*, *z* are **[EAMCET 2003]**

(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

20. Which of the following can act as an acid and as a base **[AMU 1999]**

(a) $HClO_3^-$ (b) $H_2PO_4^-$

(c) HS^- (d) All of these

21. MnO_4^{2-} (1 *mole*) in neutral aqueous medium is disproportionate to **[AIIMS 2003]**

(a) 2/3 mole of MnO_4^- and 1/3 mole of MnO_2

(b) 1/3 mole of MnO_4^- and 2/3 mole of MnO_2

(c) 1/3 mole of Mn_2O_7 and 1/3 mole of MnO_2

(d) 2/3 mole of Mn_2O_7 and 1/3 mole of MnO_2

22. The conductivity of a saturated solution of $BaSO_4$ is $3.06 \times 10^{-6}\,ohm^{-1}\,cm^{-1}$ and its equivalent conductance is $1.53\,ohm^{-1}\,cm^{-1}\,equivalent^{-1}$. The K_{sp} of the $BaSO_4$ will be **[KCET 1996]**

(a) 4×10^{-12} (b) 2.5×10^{-9}

(c) 2.5×10^{-13} (d) 4×10^{-6}

23. When MnO_2 is fused with KOH, a coloured compound is formed, the product and its colour is **[IIT Screening 2003]**

(a) K_2MnO_4, purple green

(b) $KMnO_4$, purple

(c) Mn_2O_3, brown

(d) Mn_3O_4 black

AR Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a) If both assertion and reason are true and the reason is the correct explanation of the assertion.
(b) If both assertion and reason are true but reason is not the correct explanation of the assertion.
(c) If assertion is true but reason is false.
(d) If the assertion and reason both are false.
(e) If assertion is false but reason is true.

1. Assertion : SO_2 and Cl_2 both are bleaching agents.

Reason : Both are reducing agents. **[AIIMS 1995]**

2. Assertion : Fluorine exists only in –1 oxidation state.

Reason : Fluorine has $2s^2 2p^5$ configuration. **[AIIMS 2001]**

3. Assertion : Stannous chloride is a powerful oxidising agent which oxidises mercuric chloride to mercury.

Reason : Stannous chloride gives grey precipitate with mercuric chloride, but stannic chloride does not do so. **[AIIMS 2**

4. Assertion : $HClO_4$ is a stronger acid than $HClO_3$.

Reason : Oxidation state of Cl in $HClO_4$ is +VII and in $HClO_3$ +V. **[AIIMS 2004]**

5. Assertion : In a reaction $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$, Zn is a reductant but itself get oxidized.

Reason : In a redox reaction, oxidant is reduced by accepting electrons and reductant is oxidized by losing electrons.

6. Assertion : Oxidation number of carbon in CH_2O is zero.

Reason : CH_2O formaldehyde, is a covalent compound.

7. Assertion : The oxidation numbers are artificial, they are useful as a 'book-keeping' device of electrons in reactions.

Reason : The oxidation numbers do not usually represent real charges on atoms, they are simply conventions that indicate what the maximum charge could possibly be on an atom in a molecule.

8. Assertion : H_2SO_4 cannot act as reducing agent.

Reason : Sulphur cannot increase its oxidation number beyond + 6.

9. Assertion : Equivalent weight of NH_3 in the reaction $N_2 \rightarrow NH_3$ is 17/3 while that of N_2 is 28/6.

Reason : Equivalent weight $= \frac{\text{Molecular weight}}{\text{number of } e^- \text{lost or gained}}$.

Answers

Oxidation, Reduction

1	b	2	b	3	c	4	c	5	c
6	a	7	b	8	b	9	a	10	c
11	b	12	a	13	b	14	b	15	b
16	a	17	a	18	b	19	c	20	b
21	a	22	c	23	b	24	b	25	b
26	c	27	c	28	d	29	a	30	a
31	a	32	a	33	a	34	d	35	b
36	d	37	d						

Oxidizing and Reducing agent

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

Oxidation number and Oxidation state

1	d	2	b	3	b	4	b	5	d
6	b	7	c	8	c	9	c	10	a
11	c	12	a	13	d	14	b	15	a
16	d	17	d	18	d	19	a	20	c
21	d	22	b	23	c	24	d	25	a
26	b	27	b	28	a	29	a	30	a
31	c	32	d	33	a	34	a	35	b
36	a	37	d	38	a	39	d	40	d
41	b	42	b	43	c	44	a	45	d
46	b	47	d	48	d	49	a	50	c
51	c	52	b	53	b	54	c	55	b
56	b	57	c	58	a	59	a	60	c
61	d	62	a	63	d	64	b	65	b
66	b	67	c	68	a	69	a	70	c
71	d	72	a	73	b	74	b	75	b
76	c	77	d	78	a	79	d	80	b
81	b	82	c	83	c	84	b	85	c
86	c	87	d	88	d	89	b	90	c
91	c	92	a	93	d	94	c	95	d
96	b	97	d	98	b	99	b	100	c
101	a	102	d	103	a	104	b	105	a
106	a	107	d	108	b	109	a	110	b
111	a	112	d	113	b				

Redox reaction and Method for balancing Redox reaction

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

Auto oxidation and Disproportionation

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

Critical Thinking Questions

1	d	2	c	3	a,b	4	d	5	a
6	c	7	c	8	a	9	a	10	c
11	a	12	d	13	c	14	b	15	a
16	d	17	d	18	d	19	c	20	d
21	a	22	d	23	a				

Assertion & Reason

1	c	2	b	3	e	4	b	5	a
6	b	7	a	8	a	9	a		

Answers and Solutions

Oxidation, Reduction

1. (b) $2MnO_4^{\ominus} + 5H_2O_2 + 6H^+ \rightarrow Mn^{2+} + 5O_2 + 8H_2O$.

2. (b) $S + 2e^- \rightarrow S^{2-}$

4. (c) $\overset{0}{P}_4 + 3NaOH + 3H_2O \rightarrow \underset{\text{Sodium hypophosphite}}{3NaH_2PO_2} + \overset{-3}{P}H_3$.

It shows oxidation and reduction (Redox) properties.

6. (a) In this reaction H_2S is oxidised because the oxidation state of 'S' change from – 2 to 0.

7. (b) $\overset{+4}{Pb}O_2 \rightarrow \overset{+2}{Pb}(NO_3)_2$. In this reaction reduction occurs.

8. (b) Any substance which is capable of oxidising other substances and is capable of accepting/gaining electron during oxidation is called oxidising agent or oxidant.

9. (a) $\overset{+1}{2CuI} \rightarrow \overset{O}{Cu} + \overset{+2}{CuI_2}$. Oxidation and Reduction both occur so the reaction is redox.

10. (c) $H_2S + X_2(Cl, Br, I = X) \rightarrow 2HX + S$. Here the halogen are reduced.

11. (b) When H_2O_2 reduces with $K_4[Fe(CN)_6]$. It is present in acidic solution.

$$2K_4[Fe(CN)_6 + H_2SO_4 + H_2O_2 \rightarrow$$
$$2K_3[Fe(CN)_6] + K_2SO_4 + 2H_2O$$

13. (b) In the given reaction oxidation state of *Mg* is changing from 0 to +2 while in nitrogen it is changing from 0 to –3. So oxidation of *Mg* and reduction of nitrogen takes place.

14. (b) When sodium metal is dissolved in liquid ammonia to form coloured solution. Dilute solutions are bright blue in colour due to the presence of solvated electrons.

$$Na + (x+y)NH_3 \rightarrow [Na(NH_3)_x]^+ + \underset{\text{Blue Colour}}{[e(NH_3)_y]^-}$$

15. (b) The metallic iron is oxidised to Fe^{+3}.

Oxidatio

16. (a) $\overset{+2}{SnCl_2} + \overset{+1}{2HgCl_2} \rightarrow \overset{+4}{SnCl_4} + Hg_2Cl_2(s)$

Reductio

In this reaction $HgCl_2$ is reduced in Hg.

17. (a) It is the process in which electrons are lost (de-electronation).

18. (b) $4Fe + 3O_2 \rightarrow 4Fe^{3+} + 6O^{2-}$

19. (c) *Cu* is above of *Ag* in electrochemical series and thus $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ reaction occurs.

21. (a) $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$. In this reaction Sn^{2+} change in Sn^{4+} it is called an oxidation reaction.

22. (c) $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$.

23. (b) $\overset{2+}{Zn_{(aq)}} + 2e^- \rightarrow \overset{0}{Zn_{(s)}}$ reduction.

24. (b) SO_2 bleaches by reduction while chlorine bleaches colour of flowers by oxidation.

25. (b) It is the process in which electrons are gained (electronation).

Oxidation

26. (c) $\overset{o}{Zn} + \overset{o}{I_2} \rightarrow \overset{+2\,-1}{ZnI_2}$

Reductio

In this reaction *Zn* atom oxidised to Zn^{2+} ion and iodine reduced to I^-.

27. (c) $\overset{*}{Cr}O_4^{2-}$ $\overset{*}{Cr_2}O_7^{2-}$

$x + [(-2)\times 4] = -2$ $2x + (-2)\times 7 = -2$

$x = 8 - 2 = +6$ $2x = 14 - 2 = 12$,

$x = \frac{12}{2} = +6$

In this reaction oxidation and reduction are not involved because there is no change in oxidation number.

28. (d) $3\overset{0}{Br_2} + 6CO_3^{2-} + 3H_2O \rightarrow 5\overset{-1}{Br^-} + \overset{+5}{BrO_3^-} + 6HCO_3$. In this reaction bromine is oxidised as well as reduced.

29. (a) *P* is oxidized as well as reduced (as in option a).

Reduction

30. (a) $\overset{+6}{Cr_2O_7^{2-}} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 3H_2O + 3I_2$

31. (a) In this reaction oxidation occur.

32. (a) Fluorine has highest E^o – value and more reactive than MnO_2.

33. (a) $Fe^{2+} \rightarrow Fe^{3+} + e^-$ oxidation.

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

$MnO_4^- + 5e^- \rightarrow Mn^{2+}$.

Oxidation

35. (b) $\overset{0}{Zn} + \overset{+2}{CuSO_4} \rightarrow ZnSO_4 + \overset{0}{Cu}$

Reduction

In this reaction Cu^{2+} change in Cu^o, hence it is called as reduction reaction.

36. (d) $4\overset{0}{Fe} + 3O_2 \rightarrow 4\overset{3+}{Fe} + 6O^{2-}$, in this reaction metallic iron is oxidised to Fe^{3+}.

37. (d) $2\overset{0}{N_2} + \overset{0}{O_2} \rightarrow 2\overset{+2}{N}\overset{-2}{O}$

Here O.N. of *N* increases from *O* in N_2 to +2 in *NO*, 2– and that of decreased from *O* in O_2 to –2 in *O*, therefore, it is a redox reaction.

Oxidizing and Reducing agent

1. (c) $H_2\overset{-2}{S} + H_2O_2 \rightarrow \overset{0}{S} + 2H_2O$

Oxidation

The oxidation of S shows oxidising nature of H_2O_2.

2. (a) $C_2O_4^{2-} + MnO_4^- + H^+ \rightarrow Mn^{2+} + CO_2 + H_2O$.

In this reaction $C_2O_4^{2-}$ act as a reducing agent.

3. (b) A substance which is capable of reducing other substances and is capable of donating electrons during reduction is called a reducing agent or reductant.

4. (a) Fluorine is a most powerful oxidizing agent because it consist of $E^o = +2.5\ volt$.

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

6. (b) It acts both oxidizing and reducing agent.

7. (c) Prevent action of water and salt.

9. (a) In this reaction H_2O_2 acts as a oxidizing agent.

10. (b) $NaNO_2, SnCl_2$ and HI have reducing and oxidizing properties but $NaNO_3$ have only oxidizing property.

11. (c) Because I_2 is a reducing agent.

13. (a) In this reaction H_2O acts as oxidising agent.

14. (d) I^- act as a more reducing agent than other ions.

15. (a) When sulphur dioxide is react with H_2S here SO_2 act as an oxidising agent and H_2S act as reducing agent.

16. (b) HI (Hydrogen Iodide) is a good reducing agent than other compound.

17. (b) Hydrogen sulphide (H_2S) acts as strong reducing agent as it decomposes by evolving hydrogen.

19. (b) $\overset{o}{Cl_2} + H_2O_2 \rightarrow 2H\overset{-1}{Cl} + O_2$. In this reaction chlorine reduced from zero to – 1 oxidation state.

20. (d) $NaCl + H_2O \rightarrow NaOH + HCl$

Sodium ion hydrated in water.

21. (a) Potassium has higher negative value of reduction potential hence it shows more reducing properties.

22. (b) The oxidation number of Ni changes from 0 to +1

23. (d) HNO_2 (Nitrous acid) acid acts as a oxidising, reducing agent and has complex formation properties.

24. (d) CO_2 is an oxidizing agent.

25. (b) Hydrogen peroxide (H_2O_2) act as a both oxidising and reducing agent.

27. (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 & hence it is a redox reaction.

28. (c) $Cl_2 + H_2O \longrightarrow HCl + HOCl$

$HOCl \longrightarrow HCl + [O]$

$HOCl$ can furnish, nascent oxygen.

29. (a) Reduction

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

Oxidation (reducing

30. (d) Oxidizing agent itself, undergoes reduction during a redox reaction

$H\overset{+3}{As}O_2 + \overset{+2}{Sn} \longrightarrow \overset{0}{As} + \overset{2+}{Sn} + H_2O$

Hence, here $HAsO_2$ is acting as oxidizing agent.

31. (d) $NaNO_2$ (Sodium nitrite) act both as oxidising as well as reducing agent because in it N atom is in +3 oxidation state (intermediate oxidation state)

Oxidising property

$2NaNO_2 + 2KI + 2H_2SO_4 \longrightarrow Na_2SO_4 + K_2SO_4 + 2NO + 2H_2O + I_2$

Reducing property

$H_2O_2 + NaNO_2 \longrightarrow NaNO_3 + H_2O$.

32. (c) Reduction

$\overset{0}{P} + NaOH \longrightarrow \overset{-3}{P}H_3 + NaH_2\overset{+1}{P}O_2$

Oxidation

Oxidation number and Oxidation state

1. (d) $\overset{*}{C}O_2$

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

3. (b) +2 it is a second group element.

4. (b) In $H\overset{*}{N}O_2$ oxidation number of N = + 3

In $H\overset{*}{N}O_3$ oxidation number of N = + 5.

5. (d) In case of Cl_2O chlorine shows + 1 oxidation state.

6. (b) $[\overset{*}{Cr}(H_2O)_4Cl_2]^+$

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

$x=+3$ for Cr in complex.

7. (c) $\overset{0}{Br_2} \rightarrow \overset{+5}{BrO_3^-}$, in this reaction oxidation state change from 0 to + 5.

8. (c) Oxidation state of sulphur in H_2S is –2, while it is zero in 'S' i.e. in this reaction oxidation of sulphur and reduction of chlorine is takes place.

9. (c) $K[\overset{*}{Co}(CO)_4]$

$1+x+0=0; x=-1$.

10. (a) $K_2\overset{+6}{Cr_2}O_7 \rightarrow K_2\overset{+6}{Cr}O_4$ .In this reaction no change in oxidation state of chromium.

11. (c) In hypochlorous acid chlorine atom has + 1 oxidation number.

12. (a) $S \rightarrow S^{2-}$ $O.N.$ of S = –2.

13. (d) $(NH_4)_2SO_4 \rightleftharpoons 2NH_4^+ + SO_4^{--}$

$\overset{*}{N}H_4^+$

$x+4=+1$; $x=1-4=-3$.

14. (b) In N_2O nitrogen have +1 oxidation state.

15. (a) If any central metal atom combined with corbonyl group than central metal atom shows always zero oxidation state.

16. (d) $H_2\overset{*}{S}O_4$

$2+x-2\times 4=0$, $x=8-2=+6$.

17. (d) $HCl\overset{*}{}O_4$

$1+x-2\times 4=0; 1+x-8=0$

$x=8-1=+7$ oxidation state.

18. (d) $H\overset{*}{N}O_3$; $1+x-6=0$; $x=+5$.

19. (a) Mn shows + 7 oxidation state in MnO_4^{-1}

$\overset{*}{Mn}O_4^{-1}$

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

$x-8=-1$

$x=-1+8=+7$.

20. (c) $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

21. (d) $K_2\overset{*}{Mn}O_4$

$2+x-2\times 4=0$

$x=8-2=+6$.

22. (b) Each molecule always show zero oxidation state.

23. (c) Maximum oxi. state for Cr is + 6.

24. (d) In $[Fe(CO)_5]$, transition metal Fe has zero oxidation state.

25. (a) In (b, c, d) carbon show + 4 oxidation state while in (a) carbon show – 4 oxidation state.

26. (b) $H_2\overset{*}{C_2}O_4$

$2+2x-2\times 4=0$; $2x=8-2=6$

$x=\frac{6}{2}=+3$.

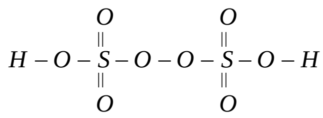
27. (b) In complex $[Pt(C_2H_4)Cl_3]^-$ Pt have + 2 oxidation state.

28. (a) $\overset{*}{C}H_2Cl_2$

$x+2-2=0$; $x=0$.

29. (a) Phosphorus shows – 3 to + 5 oxidation state.

31. (c) The chemical structure of $H_2S_2O_8$ is as follows:-



So the oxidation number of S should be :

$\underset{(\text{for } H)}{2\times(+1)} + \underset{(\text{for } S)}{2\times X} + \underset{(\text{for } O)}{6\times(-2)} + \underset{(\text{for } O-O)}{2\times(-1)} = 0$ or $X=+6$.

32. (d) In hydrazoic acid (N_3H) nitrogen shows $-\frac{1}{3}$ oxidation state.

$\overset{*}{N}_3H$

$3x+1=0$, $3x=-1$, $x=-\frac{1}{3}$.

33. (a) Hydrogen have oxidation no. + 1 and – 1.

34. (a) $[Cr(NH_3)_4Cl_2]^+$

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

$\Rightarrow x = 1 + 2 = + 3$.

35. (b) $\overset{*}{S}O_2=+4$

$H_2\overset{*}{S}O_4=+6$

$Na_2\overset{*}{S_2}O_3=+2$

$Na_2\overset{*}{S_4}O_6=+\frac{5}{2}$.

36. (a) $\overset{*}{Fe}S_2$ $\quad$ $Fe\overset{*}{S_2}$

$x-4=0$ $\quad$ $4+2x=0$

$x=+4$ $\quad$ $2x=-4$

$x=\frac{-4}{2}=-2$.

37. (d) $\overset{*}{N}O_3^-$

$x-2\times 3=-1$; $x=6-1=+5$.

38. (a) Every element always shows zero oxidation state.

39. (d) In benzaldehyde all carbon atoms show – 4 oxidation state.

40. (d) $K\overset{*}{I}O_4$

$1 + x - 2 \times 4 = 0$; $x = 8 - 1 = +7$.

41. (b) $\overset{*}{N_2}H_5^+$

$2x + 5 = +1$; $2x = 1 - 5$

$2x = -4$; $x = -2$.

42. (b) Oxidation number of C in

$HCHO = 0$

$CHCl_3 = +2$

$CH_3OH = -2$

$C_{12}H_{22}O_{11} = 0$

43. (c) $K\overset{*}{Cl}O_4$

$2 + 2x - 2 \times 7 = 0$

$2x - 14 + 2 = 0$.

44. (a) $H_4\overset{*}{I}O_6^-$

$4 + x - 12 = -1$; $x = -1 + 8 = +7$.

45. (d) Fluorine always shows – 1 oxidation state.

46. (b) In oxidation process oxidation state always increases.

47. (d) $HNO_3 \rightleftharpoons \overset{*}{N_2}O$

$1 + x - 6 = 0$ $2x - 2 = 0$

$x = +5$ $2x = 2$

$x = \frac{2}{2} = +1$.

48. (d) All free metals always shows zero oxidation state.

49. (a) $MnO_4^- \rightarrow Mn^{2+} + 5e^-$.

50. (c) C has oxidation number = 0.

51. (c) Iron has zero oxidation state in carbonyl complexes.

52. (b) In all alkali and alkaline earth metal hydride hydrogen always shows – 1 oxidation state.

53. (b) Iodine shows – 1 to + 7 oxidation state.

54. (c) Reduction

$K_2\overset{+6}{Cr_2}O_7 + 3SO_2 + H_2SO_4 \rightarrow K_2SO_4 + \overset{+3}{Cr_2}(SO_4)_3 + H_2O$

In this reaction chromium change from + 6 to +3 oxidation state.

55. (b) In H_2O_2 oxygen shows = – 1 (peroxide) oxidation state and in $BaSO_4$ oxygen shows = – 2 oxidation state.

56. (b) Mn shows highest oxidation state in $K\overset{+7}{Mn}O_4$.

57. (c) $\overset{*}{C}H_2O$

$x + 2 - 2 = 0$

$x = 0$.

58. (a) In all peroxide oxygen shows – 1 oxidation state.

59. (a) $K_2\overset{*}{Cr_2}O_7$

$2 + 2x - 2 \times 7 = 0$; $2x - 14 + 2 = 0$

$2x = 12$; $x = \frac{12}{2} = +6$.

60. (c) Nickle shows zero oxidation state in carbonyl complex.

61. (d) $\overset{*}{Os}O_4$

$x + 4(-2) = 0$

$x - 8 = 0$

$x = +8$.

62. (a) Al shows + 3 oxidation state.

63. (d) $K_4[\overset{*}{Fe}(CN)_6]$

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

$x = +2$.

In this complex compound Iron show + 2 oxidation state.

64. (b) In this complex iron is a central metal atom showing + 2 oxidation state.

65. (b) Oxygen shows + 2 oxidation state in F_2O . As F most electronegative element, it always has an O. No. =–1

66. (b) $H_3\overset{*}{P}O_3$

$3 + x - 2 \times 3 = 0$; $x = 6 - 3 = +3$.

67. (c) $Mg_2\overset{*}{P_2}O_7$

$4 + 2x - 2 \times 7 = 0$; $2x = 14 - 4 = 10$

$2x = 10$; $x = \frac{10}{2} = +5$.

68. (a) $3 \times x + 1(1) = 0$

$3x + 1 = 0$

$3x = -1, \Rightarrow x = -\frac{1}{3}$ in N_3H

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

$x = -1$ in NH_2OH

$x \times 2 + 4(1) = 0$ $x = -\frac{4}{2} = -2$ in N_2H_4

$x + 3(1) = 0$ $x = -3$ in NH_3

Hence, highest in N_3H .

69. (a) In $KH_2\overset{*}{P}O_2$

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

$3 + x - 4 = 0$; $x = +1$.

70. (c) Oxygen has 6 electrons in the outer most shell and shows common oxidation state – 2.

71. (d) $H_2\overset{*}{S}O_3 = +4$; $\overset{*}{S}O_2 = +4$

$H_2\overset{*}{S}O_4 = +6$; $H_2\overset{*}{S} = -2$.

72. (a) The oxidation number of sulphur in the sulphur molecule (S_8) is 0 and 2.

73. (b) In ferrous ammonium sulphate Fe shows +2 oxidation state.

74. (b) $\overset{*}{N}H_2OH$

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

$x + 2 - 2 + 1 = 0$; $x = -1$.

75. (b) $Ba(H_2\overset{*}{P}O_2)_2$; $BaH_4\overset{*}{P}_2O_4$

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

$x = \frac{2}{2} = +1$.

77. (d) $H_2\overset{*}{S}O_4$

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

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

Electronic configuration of sulphur in H_2SO_4 is

$1s^2, 2s^2, 2p^6$.

78. (a) $K\overset{*}{Mn}O_4$

$1 + x - 2 \times 4 = 0$; $x = 8 - 1 = +7$.

79. (d) H_3AsO_4

$+3 + x - 2 \times 4 = 0$; $x = 8 - 3 = +5$.

80. (b) The oxidation state of Xe in both XeO_3 and XeF_6 is + 6

$\overset{*}{X}eO_3$ $\quad$ $\overset{*}{X}eF_6$

$x - 2 \times 3 = 0$ $\quad$ $x - 6 = 0$

$x = +6$ $\quad$ $x = +6$.

81. (b) $\overset{*}{C}H_3 - Cl$

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

$x + 3 - 1 = 0$; $x + 2 = 0$

$x = -2$.

82. (c) $\overset{*}{Cr}_2O_7^{2-}$

$2x - 2 \times 7 = -2$; $2x = 14 - 2 = 12$

$x = \frac{12}{2} = +6$.

83. (c) $H_2\overset{*}{S}O_3$

$+2 + x - 2 \times 3 = 0$; $x = 6 - 2 = +4$.

84. (b) Two Cl atom shows +1 and -1 oxidation state.

85. (c) $H\overset{*}{Cl}O_3$

$1 + x - 2 \times 3 = 0$; $x = 6 - 1 = +5$.

86. (c) $5 \begin{array}{c} COOH \\ | \\ COOH \end{array} + 2K\overset{+7}{Mn}O_4 + 3H_2SO_4 \rightarrow$

$K_2SO_4 + 2\overset{+2}{Mn}SO_4 + 10CO_2 + 8H_2O$

In this reaction oxidation state of Mn change from + 7 to + 2.

87. (d) Oxygen have + 2 oxidation state in OF_2 .

89. (b) $\overset{*}{S}_2O_3^{2-}$

$2x + 3(-2) = -2$; $x = +2$.

90. (c) $x + 2 \times (+1) + 2(-1) = 0$

$x + 2 - 2 = 0$; $x = 0$ in CH_2Cl_2 .

91. (c) In potassium superoxide (KO_2) oxygen shows, $-\frac{1}{2}$ oxidation state.

92. (a) $\overset{*}{S}_2Cl_2$

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

$x = +1$.

93. (d) $Na_2\overset{*}{S}_4O_6$

$2 + 4x - 12 = 0$

$4x = 10$ $x = \frac{10}{4}$ $x = \frac{5}{2}$.

94. (c) $\overset{+2}{Cu}SO_4 + 2KI \rightleftharpoons K_2SO_4 + CuI_2$

$2CuI_2 \longrightarrow \overset{+1}{Cu_2}I_2 + I_2$

95. (d) $NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$

$\overset{*}{N}H_4^+$

$x + 4 = +1$; $x = 1 - 4 = -3$.

96. (b) $2\overset{+4}{NO_2} + H_2O \rightarrow HNO_2 + H\overset{+5}{N}O_3$. In this reaction oxidation state changes.

97. (d) $\overset{*}{Fe}_3O_4$

$3x + (-8) = 0$; $3x - 8 = 0$

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

99. (b) $K_2\overset{*}{Mn}O_4$ $\quad$ $\overset{*}{Mn}SO_4$

$2 + x - 8 = 0$ $\quad$ $x + 6 - 8 = 0$

$x = +6$ $\quad$ $x = +2$.

100. (c) Chlorine have oxidation state – 1 to + 7.

101. (a) $[\overset{*}{Co}(NH_3)_4ClNO_2]$

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

$x + 0 - 1 - 1 = 0$

$x - 2 = 0$; $x = +2$.

102. (d) $K_4[\overset{*}{Ni}(CN)_4]$

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

$+4 + x - 4 = 0 \Rightarrow x = 0$.

103. (a) Fluorine always shows – 1 oxidation state in oxides.

104. (b) $K_3[\overset{*}{Fe}(CN)_6]$

$1 \times 3 + x + (-1 \times 6) = 0$

$3 + x - 6 = 0$; $x = +3$.

105. (a) $\overset{*}{N}H_3$

$x + 3(+1) = 0$, $x = -3$.

106. (a) $_{26}Fe \longrightarrow [Ar]3d^6 4S^2$

$Fe^{++} \longrightarrow [Ar]3d^6 4S^0$

$Fe^{+++} \longrightarrow [Ar]3d^5 4S^0$

In +2 state Fe is called Ferrous & in +3 state as ferric.

107. (d) Fluorine is the most electronegative element in the periodic table so it never shows positive oxidation state.

108. (b) Silicon forms silicides with strongly electropositive metals (like Na, Mg, K etc.) In these compounds. It has oxidation number $= -4$.

109. (a) H_2S [O.N. of $H = +1$]

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

$2 + x = 0$; $x = -2$

110. (b) Let the oxidation number of N in $NaNO_2$ be x

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

$1 + x - 4 = 0$; $x = +3$

111. (a) $x = 8 - 2 = +6$

112. (d) $K_2Cr_2O_7 + 6KI + 7H_2SO_4 \rightarrow 4K_2SO_4 + Cr_2(SO_4)_3$ $+ 7H_2O + 3I_2$

$\overset{*}{C}r_2(SO_4)_3 \rightarrow 2\overset{+3}{Cr} + 3SO_4^{2-}$

113. (b) Let the oxidation number of I in $IPO_4 = x$

Oxidation number of $PO_4 = -3$

$x + (-3) = 0 \Rightarrow x = +3$

Redox reaction and Method for balancing Redox reaction

1. (a) $MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{++} + 4H_2O$.

2. (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.

3. (c) $\underbrace{Cr_2O_7^{2-} + 3e^- \rightarrow Cr^{3+}}_{\text{Reduction}}$.

In this reaction three electrons are required for the reduction of $Cr_2O_7^{2-}$ into Cr^{3+}.

4. (c) Number of e^- transferred in each case is 1, 3, 4, 5.

5. (a) Starch paper are used for iodine test

as: $I^- + \text{oxidant} \longrightarrow I_2$

$I_2 + \text{starch} \longrightarrow \text{blue colour}$

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

$(Sn^{2+} \rightarrow Sn^{4+} + 2e^-) \times 3$

$Cr_2O_7^{2-} + 14H^+ + 3Sn^{2+} \rightarrow 3Sn^{4+} + 2Cr^{3+} + 7H_2O$

It is clear from this equation that 3 *moles* of Sn^{2+} reduce one mole of $Cr_2O_7^{2-}$, hence 1 *mol*. of Sn^{2+} will reduce $\frac{1}{3}$ *moles* of $Cr_2O_7^{2-}$.

7. (a) $2MnO_4^{\ominus} + 5H_2O_2 + 6H^+ \rightarrow 2Mn^{2+} + 5O_2 + 8H_2O$.

8. (b) $2Fe^{3+} + Sn^{2+} \rightarrow 2Fe^{2+} + Sn^{4+}$ (Reduction: $Fe^{3+} \rightarrow Fe^{2+}$; Oxidation: $Sn^{2+} \rightarrow Sn^{4+}$)

9. (a) $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O \times 2$

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

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

Thus the coefficient of MnO_4^-, $C_2O_4^{2-}$ and H^+ in the above balanced equation respectively are 2, 5, 16.

10. (d). $\overset{0}{Zn} + 2\overset{+1}{Ag}CN \rightarrow 2\overset{0}{Ag} + \overset{+2}{Zn}(CN)_2$. (Reductio; Oxidatio)

11. (d) $\overset{o}{Cu} + 2\overset{+1}{Ag}NO_3 \rightarrow \overset{+2}{Cu}(NO_3)_2 + 2\overset{o}{Ag}$. This is a redox reaction. (Oxidation; Reduction)

12. (c) $\overset{0}{H_2} + \overset{0}{Br_2} \rightarrow 2\overset{+}{H} - \overset{-}{Br}$ (Oxidatio; Reductio)

13. (c) Higher is the reduction potential stronger is the oxidising agent. Hence in the given options. MnO_4^- is strongest oxidising agent.

14. (a) $IO_3^- + aI^- + bH^+ \rightarrow cH_2O + dI_2$

Step 1 : $I^{-1} \rightarrow I_2$ (oxidation)

$IO_3^- \rightarrow I_2$ (reduction)

Step 2 : $2IO_3^- + 12H^+ \rightarrow I_2 + 6H_2O$

Step 3 : $2IO_3^- + 12H^+ + 10e \rightarrow I_2 + 6H_2O$

$2I^- \rightarrow I_2 + 2e$

Step 4 : $2IO_3^- + 12H^+ + 10e^- \rightarrow I_2 + 6H_2O$

$[2I^- \rightarrow I_2 + 2e]5$

Step 5 : $2IO_3^- + 10I^- + 12H^+ \rightarrow 6I_2 + 6H_2O$

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

On comparing, $a = 5$, $b = 6$, $c = 3$, $d = 3$

15. (d) *In alkaline medium*

$2KMnO_4 + KI + H_2O \rightarrow 2MnO_2 + 2KOH + KIO_3$.

Auto oxidation and Disproportionation

1. (d) $H_2S \rightarrow \overset{0}{S} + 2e$

Equivalent wt. = $\frac{\text{Mol.wt.}}{2} = \frac{34}{2} = 17$.

2. (a) $1.12\, ltr\, H_2 = 1.2\, g; \therefore 22.4\, ltr\, H_2 = 24\, g$.

3. (d) $2AgNO_3 \xrightarrow{\Delta} 2Ag + 2NO_2 + O_2$.

4. (b) To prevent rancidification of food material we add anti-oxidant which are called oxidation inhibitor.

6. (b) $Zn^{2+}/Zn.E^o = -0.76\, V$

$Al^{3+}/Al \quad E^o = -1.662$

$Sn^{2+}/Sn \quad E^o = -0.136$

$Pb^{2+}/Pb \quad E^o = -0.126$

In galvanizing action Zn is coated over iron.

8. (d) Molecular weight of H_3PO_4 is 98 and change in

its valency = 1 equivalent wt. of H_3PO_4

$$= \frac{\text{Molecular weight}}{\text{Change in valency}} = \frac{98}{1} = 98.$$

9. (b) Equivalent mass

$$= \frac{\text{Molecular weight}}{\text{Change in oxidation number per mole}}$$

Suppose molecular weight is M

Oxidation number of I_2 in IO_4^- in

Acidic medium i.e., $I \times (-8) + 1e^- = +7$

So eq. wt. $= M/7$.

10. (c) $2KMnO_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5O$

$5H_2O_2 + 5O \longrightarrow 5H_2O + 5O_2$

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

11. (a) $\frac{\text{Molecular weight}}{2}$ = Equivalent weight of Iodine.

12. (d) $\frac{\text{Molecular weight}}{3}$ Because in KIO_3 effective oxidation number is 3.

13. (a) $6MnO_4^- + I^- + 6OH^- \longrightarrow 6MnO_4^{2-} + IO_3^- + 3H_2O$

14. (c) $ClO_2 \rightarrow Cl^-$

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

$H_2O_2 \rightarrow O_2$

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

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

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

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

$2ClO_2 \equiv 5H_2O_2$

$\therefore \quad ClO_2 = 2.5H_2O_2$

Critical Thinking Questions

1. (d) HNO_2 shows both oxidation and reduction properties.

2. (c) Al_2O_3 could not act as a oxidising and reducing agent.

3. (a, b) In H_2S sulphur shows –2 oxidation state and in SO_2 shows +4 oxidation state. Hence SO_2 shows both oxidising and reducing properties.

4. (d) All the given statements are true.

5. (a) $H-O-\overset{\overset{H}{|}}{\underset{\underset{O}{\downarrow}}{P}}-OH$, hence it is dibasic. It acts as reducing agent also.

6. (c) (a) $\overset{*}{N}O_2$; $x-4=0$; $x=+4$

(b) $H\overset{*}{N}O$; $1+x-2=0$; $x=+1$

(c) $\overset{*}{N}H_3$; $x+3=0$; $x=-3$

(d) $\overset{*}{N}_2O_5$; $2x-10=0$; $2x=10$; $x=\frac{10}{2}$; x = 5.

7. (c) 2 × No. of e^- losses = Oxi. no.

$2\times 3e^-=+6$.

8. (a) The ion which is not affected during the course of reaction is known as spectator ion.

9. (a) $H_2\overset{*}{S}_2O_7$

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

$+2+2x-14=0$

$2x=14-2=12$

$x=\frac{12}{2}=+6$ for S

$K_4\overset{*}{Fe}(CN)_6$

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

$4+x-6=0$

$x=6-4=+2$ for Fe .

10. (c) $K\overset{*}{O}_2$, $+1+2x=0,\ x=-\frac{1}{2}$.

11. (a) $N_2^{2-}\rightarrow {}_2N^{a+}+10e^-$

$\therefore\ 2a-[2\times(-2)]=10$

$\therefore\ a=+3$.

12. (d) $\overset{*}{Cr}O_2Cl_2$, $x-4-2=0, x=+6$.

13. (c) $3x=-1,\ x=-1/3$.

14. (b) $Ba_2\overset{*}{Cu}_3O_7$

$3+2\times 2+3x-(2\times 7)=0$

$3+4+3x-14=0$

$3x=7$

$x=\frac{7}{3}$.

15. (a) $\overset{*}{S}_8=0$

$\overset{*}{S}_2F_2=+1$

$H_2\overset{*}{S}=-2$.

16. (d) In reaction $HCl+H_2O\rightarrow H_3O^-+Cl^-$, only reduction has taken place not oxidation.

17. (d) Zn can oxidise carbon because heat of combustion of $Zn<C$.

18. (d) $B_2H_6+2KOH+2H_2O\rightarrow 2KBO_2+6H_2$.

19. (c) The values of x,y,z are 8, 4, 4 respectively hence the reaction is

$H_2SO_4+8HI\rightarrow H_2S+4I_2+4H_2O$

20. (d)

Acid	Base
$HClO_3^-$	ClO_3^{2-}
HS^-	S^{2-}
$H_2PO_4^-$	HPO_4^{2-}

21. (a) MnO_4^{2-} in neutral aqueous medium is disproportionate to $\frac{2}{3}$ mole of MnO_4^- and $\frac{1}{3}$ mole of MnO_2 .

22. (d) $\lambda m=\frac{1000\ K}{S}=\frac{1000\times 3.06\times 10^{-6}}{S}=1.53$

$S=2\times 10^{-3}\frac{mol}{litre}$

$K_{sp(BaSO_4)}=S^2=(2\times 10^{-3})^2=4\times 10^{-6}$.

23. (a) $2MnO_2+4KOH+O_2\xrightarrow{\Delta}\underset{\text{purple green}}{2K_2MnO_4}+2H_2O$.

Assertion & Reason

1. (c) It is true that SO_2 and Cl_2 both are bleaching agents. But Cl_2 is an oxidising agent while SO_2 is a reducing agent. Therefore, in this questions assertion is true while reason is false.

2. (b) It is correct that fluorine exists only in –1 oxidation state because it has $1s^2 2p^5$ electronic configuration and thus shows only –1 oxidation state in order to complete its octet. Hence, both assertion and reason are true and reason is not a correct explanation of assertion.

3. (e) Here, assertion is false, because stannous chloiride is a strong reducing agent not strong oxidising agent. Stannous chlorides gives Grey precipitate with mercuric chloride. Hence, reason is true.

4. (b) Both assertion and reason are true but reason is not the correct explanation of assertion. Greater the number of negative atoms present in the oxy-acid make the acid stronger. In general, the strengths of acids that have general formula $(HO)_m ZO_n$ can be related to the value of n. As the value of n increases, acidic character also increases. The negative atoms draw electrons away from the Z-atom and make it more positive. The Z-atom, therefore, becomes more effective in with drawing electron density away from the oxygen atom that bonded to hydrogen. in turn, the electrons of $H-O$ bond are drawn more strongly away from the H-atom. The net effect makes it easier from the proton release and increases the acid strength.

5. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

Oxidation loss of 2*e*

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

Reduction gain of 2*e*

6. (b) Both assertion and reason are true but reason is not the correct explanation of assertion.

Oxidation number can be calculated using some rules. H is assigned +1 oxidation state and O has oxidation number –2

$\therefore$ O. No. of C in CH_2O:

O. no. of $C + 2(+1) + (-2) = 0$

$\therefore$ O. No. of $C = 0$

7. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

8. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

Maximum oxidation state of S is +6, it cannot exceed it. Therefore it can't be further oxidised as S^{-2} can't be reduced further.

9. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

$$\overset{0}{N_2} + 6e^- \longrightarrow 2N^{3-}$$

$\therefore$ equivalent weight of

$NH_3 = \frac{14+3}{3} = \frac{17}{3}$ (M. wt. of NH_3)

while for $N_2 = \frac{14 \times 2}{6} = \frac{28}{6}$

Redox Reactions

SET Self Evaluation Test - 13

1. When a piece of wire of copper is dipped in $AgNO_3$ solution, the colour of the solution turns blue due to **[MP PMT 1992; JIPMER 2002]**
   (a) Formation of soluble complex
   (b) Oxidation of copper
   (c) Oxidation of silver
   (d) Reduction of copper

2. HBr and HI can reduce H_2SO_4, HCl can reduce $KMnO_4$ and HF can reduce **[IIT 1981]**
   (a) H_2SO_4 (b) $KMnO_4$
   (c) $K_2Cr_2O_7$ (d) None of the above

3. Consider the following statements :
   In the chemical reaction
   $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$
   (1) Manganese ion is oxidised
   (2) Manganese ion is reduced
   (3) Chloride ion is oxidised
   (4) Chloride ion is reduced
   Which of these statements are correct **[NDA 1999]**
   (a) 1 and 3 (b) 1 and 4
   (c) 2 and 3 (d) 2 and 4

4. The oxide which cannot act as a reducing agent is **[CBSE PMT 1995; AIIMS 2000; JIPMER 2002; Kurukshetra CEE 2002]**
   (a) SO_2 (b) NO_2
   (c) CO_2 (d) ClO_2

5. In the reaction between ozone and hydrogen peroxide, H_2O_2 acts as **[RPET 2000]**
   (a) Oxidising agent
   (b) Reducing agent
   (c) Bleaching agent
   (d) Both oxidising and bleaching agent

6. The oxidation state of each oxygen atom in Na_2O_2 is **[NCERT 1971]**
   (a) – 2 each (b) – 2 and zero
   (c) – 1 each (d) None of the above

7. The oxidation state of sulphur in SO_4^{2-} is **[Bihar MEE 1996]**
   (a) 4 (b) 2
   (c) 6 (d) – 6

8. The charge on cobalt in $[Co(CN)_6]^{3-}$ is **[CPMT 1985, 93]**
   (a) – 6 (b) – 3
   (c) + 3 (d) + 6

9. Oxidation number of S in Na_2SO_4 is **[CPMT 1989]**
   (a) – 2 (b) + 2
   (c) – 6 (d) + 6

10. A metal ion M^{3+} after loss of three electrons in a reaction will have an oxidation number equal to **[CPMT 1980, 83, 84, 94, 99]**
    (a) Zero (b) + 2
    (c) + 3 (d) + 6

11. Oxidation number of oxygen in ozone (O_3) is **[MP PET 2000; MP PMT 2001]**
    (a) + 3 (b) – 3
    (c) – 2 (d) 0

12. The oxidation states of sulphur in the anions $SO_3^{2-}, S_2O_4^{2-}$ and $S_2O_6^{2-}$ follow the order **[CBSE PMT 2003]**
    (a) $S_2O_6^{2-} < S_2O_4^2 < SO_3^{2-}$ (b) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$
    (c) $SO_3^{2-} < S_2O_4^{2-} < S_2O_6^{2-}$ (d) $S_2O_4^2 < S_2O_6^{2-} < SO_3^{2-}$

13. The oxidation number of hydrogen in LiH is
    (a) + 1 (b) – 1
    (c) 2 (d) 0

14. Which of the following is not a redox reaction **[RPMT 1999]**
    (a) $2Rb + 2H_2O \rightarrow 2RbOH + H_2$
    (b) $2CuI_2 \rightarrow 2CuI + I_2$
    (c) $2H_2O_2 \rightarrow 2H_2O + O_2$
    (d) $4KCN + Fe(CN)_2 \rightarrow K_4Fe(CN)_6$

15. Which of the following equations is a balanced one **[EAMCET 1980]**
    (a) $5BiO_3^- + 22H^+ + Mn^{2+} \rightarrow 5Bi^{3+} + 7H_2O + MnO_4^-$
    (b) $5BiO_3^- + 14H^+ + 2Mn^{2+} \rightarrow 5Bi^{3+} + 7H_2O + 2MnO_4^-$
    (c) $2BiO_3^- + 4H^+ + Mn^{2+} \rightarrow 2Bi^{3+} + 2H_2O + MnO_4^-$
    (d) $6BiO_3^- + 12H^+ + 3Mn^{2+} \rightarrow 6Bi^{3+} + 6H_2O + 3MnO_4^-$

16. In the equation
    $4M + 8CN^- + 2H_2O + O_2 \rightarrow 4[M(CN)_2]^- + 4OH^-$
    Identify the metal M **[AFMC 1998]**
    (a) Copper (b) Iron
    (c) Gold (d) Zinc

17. In alkaline condition $KMnO_4$ reacts as $2KMnO_4 + 2KOH \rightarrow 2K_2MnO_4 + H_2O + O$. The equivalent weight of $KMnO_4$ would be (Atomic mass of K = 39, Mn = 55, O = 16) **[MP PMT 2002]**
    (a) 158.0 (b) 79.0
    (c) 52.7 (d) 31.6

18. In acidic medium, equivalent weight of $K_2Cr_2O_7$ $(mol.\ wt. = M)$ is **[AFMC 1988]**

(a) $M/3$ (b) $M/4$ (c) $M/6$ (d) $M/2$

Answers and Solutions

(SET -13)

1. (b) $2Ag^+ + Cu \rightarrow Cu^{++} + 2Ag^-$; $E^o_{Ag^+/Ag} > E^o_{Cu^{++}/Cu}$.

2. (d) F^- can be oxidised to F_2 only by electrolysis.

3. (c) Because the oxidation state of chlorine is – 4 to 0 while Manganese ion is reduced because its oxidation state + 4 to + 2.

4. (c) CO_2 is a acidic oxide.

5. (b) H_2O_2 acts as a reducing agent in the reaction between O_3 and H_2O_2.

6. (c) In Na_2O_2 oxygen show – 1 oxidation state.

7. (c) $\overset{*}{S}O_4^{2-}$

$x - 2 \times 4 = -2$

$x = 8 - 2 = +6$.

8. (c) In $[Co(CN)_6]^{3-}$ complex Co shows + 3 oxidation state.

9. (d) $Na_2\overset{*}{S}O_4$

$2 + x - 2 \times 4 = 0$

$x = +6$.

10. (d) $M^{3+} \rightarrow M^{6+} + 3e^-$. Thus the oxidation number of metal = + 6.

11. (d) Molecule and free atoms show zero oxidation state O_3 is a molecule shows zero oxidation state.

12. (b) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$

Oxi. state of sulphur in $S_2O_4^{2-} = +3$

Oxi. state of sulphur in $SO_3^{2-} = +4$

Oxi state of sulphur in $S_2O_6^{2-} = +5$.

13. (b) $\overset{+1}{Li}\overset{-1}{H}$.

14. (d) In the reaction $4KCN + Fe(CN)_2 \rightarrow K_4Fe(CN)_6$, change in oxidation state is not taking place.

15. (b) $5BiO_3^- + 14H^+ + 2Mn^{2+} \rightarrow 5Bi^{3+} + 7H_2O + 2MnO_4^-$ is the balanced reaction.

16. (c) $4Au + 8CN^- + 2H_2O + O_2 \rightarrow 4[Au(CN)_2]^- + 4OH^-$.

17. (a) $e^- + Mn^{7+} \rightarrow Mn^{6+}$ $\therefore E = \frac{M}{1}$.

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

Equivalent weight of $K_2Cr_2O_7$

$= \frac{\text{Molecular Mass}}{6} = \frac{294.2}{6} = \frac{M}{6}$.

***