CHEMISTRY

MATRICULATION

10

Untouchability is a sin
Untouchability is a crime
Untouchability is inhuman

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PREFACE

What is chemistry chemistry is ‘Pulling substances apart to find out what they are made of and putting things together to make new substances’

We have great pleasure in presenting a textbook in chemistry for class X Matriculation students. The book has been prepared strictly according to the new syllabus prescribed by the Directorate of Matriculation Schools, Government of Tamilnadu.

The important features of the book are

- The subject matter is written in a simple manner, with suitable examples cited wherever required.
- A large number of solved problems are introduced at the appropriate places, to enable the student understand the underlying principle and laws clearly.
- Self-evaluation covering short and lengthy questions and unsolved problems are given at the end of each chapter.
- Practical syllabus is also provided in the same way as needed to conduct the practical in the laboratory.

I sincerely hope that the present book will be appreciated by my leaned colleagues and the young reders. Suggestions for the improvement of the present edition will be highly appreciated and incormporated in the next edition.

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CONTENTS

1. ATOMIC STRUCTURE

INTRODUCTION:

All objects around us, this book, your pen, and other things of nature such as rocks, water plants and animal substances – constitute the matter of the universe. Matter can be classified into pure substances and mixtures. A pure substance is a single, uncontaminated substance. A mixture is a physical combination of two or more pure substances. There are two types of pure substances known as ***elements and compounds***.
Elements are fundamental substances, which cannot be fragmented into simpler fundamental substances. Compounds are composed of elements and they can be separated into its constituent elements. The British chemist John Dalton (1766-1844) provided the basic theory that *all matter is* ***composed of small particles called atoms***.

In 1897, the British physicist J.J. Thomson discovered the first component part of the atom *the electron* and in 1904 he proposed and initial model of an atom. In 1911, Rutherford put forward the idea of the nuclear model of the atom, based on experiments done in his laboratory. But he could not explain the stability of atom. In order to take account of atomic stability, in 1913 Niels Bohr, created a new model of atom.

Until 1923, all attempts to deal with atomic and molecular structural problems were based on ***classical mechanics,*** in which structural units of the atom were treated as *particles.* The modern theory of atomic structure is based o the *quantum* or wave *mechanics* proposed independently by de Broglie, Heisenberg and Schrodinger.

IMPORTANT TERMS, DEFINITIONS AND FORMULAE

ATOM: An atom is the smallest particle of matter consisting of a positively charged nucleus and negatively charged electrons.

DALTON'S ATOMIC MODEL: This model suggest that atoms are indivisible.

J.J. THOMSON'S MODEL: An atom is a solid sphere of positively charged particles in which electrons are embedded like seeds in watermelon fruit.

RUTHERFORD'S MODEL. The protons and neutrons are present in a small dense positively charged core called nucleus and all the electrons revolve around the nucleus in circular paths like planetary model.

BOHR'S MODEL. The protons and neutrons are present in the nucleus and all the electrons revolve around the nucleus in definite orbits.

SOMMERFIELD MODEL: According to this, the electron moving around the nucleus must describe an elliptical orbit in addition to circular orbits as suggested by Bohr.

QUANTUM MECHANICAL MODEL OF AN ATOM: According to this, electrons are considered as **three dimensional wave** in electric field of the positively charged nucleus.

ELECTRON: It is a negatively charged particle which occupies the space outside the nucleus in an atom.

PROTON: It is a positively charged particles present in the nucleus of an atom.

NEUTRON: It is a neutral particle of mass equal to the mass of proton.

ORBIT: It is a definite circular path in which the electron is supposed to revolve around the nucleus.

ORBITAL: It is the three dimensional region around the nucleus in which the probability of finding the electron is maximum.

ELECTRONIC CONFIGURATION: Distribution of electron in different orbitals of the atom of an element.

s-ORBITAL: s orbital is spherically symmetric around the nucleus.

p-OTBITAL: p orbital is dumb-bell shaped and consists of two lobes of electron cloud.

PRINCIPAL QUANTUM NUMBER (*n*): It describes the energy of the energy level in which the electron revolving around the nucleus. It also describes the distance between the nucleus and the electron.

AZIMUTHAL OR ORBITAL QUANTUM NUMBER (*l*): It represents the sub energy level which is present in main energy level.

MAGNETIC QUANTUM NUMBER (*m*): It gives how many orientations are possible for a sub energy level in space, when an electron present in a sub shell.
SPIN QUANTUM NUMBER (*s*): It describes the direction of the spin of the electron (either clock wise or anticlock wise).

HUND'S RULE: "Among the orbitals of same energy, electrons do not start pairing, until all these orbitals are singly occupied". Hund's rule is also called as the principle of minimum pairing and the principle of maximum multiplicity.

AUFBAU'S PRINCIPLE: "Electrons are filled in the increasing order of energy level" According to this principle first the electrons occupy the orbitals with lowest energy. This is decided by the sum of the principle quantum number and azimuthal quantum number. This is called ($n + l$) rule.

PAULI PRINCIPLE: "In an atom no two electrons can have the same set of four quantum numbers".

OXIDATION: Removal of one or more electron from an atom or molecule or ion is called oxidation

REDUCTION: Addition of one or more electrons to an atom or molecule or ion is called reduction.

REDUCING AGENT: A substance which gives one or more electrons to the other is called a reducing agent.

OXIDISING AGENT: A substance which accepts one or more electrons from the other is called an oxidizing agent.

SELF EVALUATION (T.B. PAGE 19 & 20)

I. CHOOSE THE CORRECT ANSWER:

1. What would be the *1* value for an electron having n value 2?
   a) 1 b) 2 c) 2, 0 d) 0, 1
2. The quantum number which gives the orientation of a given electron is
   a) Principal quantum number b) Magnetic quantum number
   c) Azimuthal quantum number d) Spin quantum number
3.In an atom no two electrons have the same set of four quantum numbers. It is
   a) Pauli's exclusion principle b) Hund's rule c) Aufbau principle d) Bohr's principle
4.The ground state electronic configuration of carbon is
   a) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$ b) $1\ s^2\ 2s^1 2p_x^1\ 2p_y^1\ 2p_z^1$
   c) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^0\ 2p_z^0$ d) $1s^1\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$
5.The oxidation number of manganese in $KMnO_4$ is
   a) +2 b) –2 c) +1 d) +7
6. The direction of the spin of electron is explained by ………..quantum number.
   a) Principal b) Azumithal c) magnetic d) Spin
7. The quantum number which gives the maximum number of electrons in a shell is……………. quantum number
   a) Principal b) Azumithal c) Magnetic d) Spin
8. The electron will enter subshell for which (n + *l*) value is………..
   a) Lowest b) Highest c) both (a) & (b) d) None
9. Oxidation number of hydrogen in LiH is…………
   a) +1 b) –1 c) +2 d) –2
10.A substance which accepts one or more electrons from the other is called ………..
   a) Reducing agent b) Oxidising agent c) Redox reagent d) None

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

II. ANSWER THE FOLLOWING IN ONE OR TWO SENTENCES: (T.B PAGE 21)

1.State Hund's rule.

"Among the orbitals of same energy, electrons do not start pairing, until all these orbitals are singly occupied".

Hund's rule is also called as the principle of minimum pairing and the principle of maximum multiplicity.

2.Define oxidation and reduction in terms of electron transfer.

Removal of one or more electrons from a particle is called **oxidation**. Addition of one or more electrons to a particle is called **reduction**.

Example,

1. $Na \longrightarrow Na^+ + e^-$ (oxidation)
2. $Cl + e^- \longrightarrow Cl^-$ (Reduction)

3. Draw the shape of *p*-orbitals.

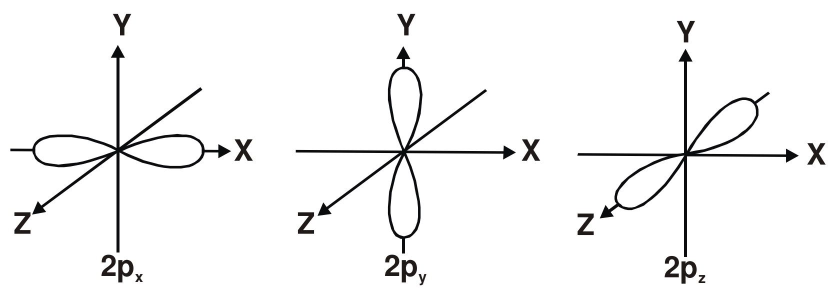

Fig. Shapes of $2p_x$, $2p_y$ and $2p_z$ orbitals

4.What do you mean by oxidation number?

Oxidation number is defined as the actual or apparent charge possessed by an atom of the element in a compound.

5.Calculate the oxidation number of

(a) Cr in $K_2Cr_2O_7$

Solution: Oxidation of K is +1, and oxygen is – 2.

$$\begin{aligned} \therefore\ 2(1) + 2Cr + 7 \times -2 &= 0 \\ 2 + 2Cr - 14 &= 0 \\ 2Cr - 12 &= 0 \\ 2Cr &= +12 \\ Cr &= \frac{+12}{2} \\ \mathbf{Cr} &= \mathbf{+6} \end{aligned}$$

∴ The oxidation number of Cr in $K_2Cr_2O_7$ is +6

(b) Cl in ClO_3^-

Solution: Oxidation number of oxygen is – 2.

$$\begin{aligned} Cl\,O_3 &= -1 \\ Cl + 3 \times -2 &= -1 \\ Cl - 6 &= -1 \\ Cl &= -1 + 6 \\ \mathbf{Cl} &= \mathbf{+5} \end{aligned}$$

∴ The oxidation number of Cl in ClO_3^- is +5

(c) S in $Na_2S_2O_3$

Solution: Oxidation number of Na is +1, Oxygen is –2.

$$\begin{aligned} Na_2S_2O_3 &= 0 \\ 2 \times 1 + 2S + 3 \times -2 &= 0 \end{aligned}$$

$$2 + 2S - 6 = 0$$
$$2S - 4 = 0$$
$$2S = +4$$
$$S = 4/2$$

S = +2

∴ The oxidation number of S in $Na_2S_2O_3$ is +2

III. ANSWER IN BRIEF (T.B. Page 21)

1.State Aufbau's principle.

"Electrons are filled in the increasing order of energy level"

According to this principle first the electrons occupy the orbitals with lowest energy. This is decided by the sum of the principle quantum number and azimuthal quantum number. This is called ($n + l$) rule.

Rule 1: The electrons first occupy that orbital for which ($n + l$) value is lowest.

Rule 2: When ($n + l$) values for two orbitals are equal, then the electrons first occupy the orbital with lower value of n.

2.State Pauli exclusion principle with an illustration.

"In an atom no two electrons can have the same set of four quantum numbers".

Illustration of Pauli's exclusion principle:

1. In an atom if one electron is assigned a set of four quantum numbers $n = 1$, $l = 0$, $m = 0$, $s = +½$, then other electrons **cannot be assigned** the same set of quantum numbers.
2. If three quantum numbers for two electrons are the same, then these electrons must have different fourth quantum number.

	n	l	m	s
First electron	1	0	0	+ ½
Second electron	1	0	0	– ½

3.Give the difference between orbit and orbital.

Difference between an orbit and an orbital

ORBIT	ORBITAL
1. It is a definite circular path in which the electron is supposed to revolve around the nucleus.	It is the three dimensional region around the nucleus in which there is maximum probability of finding the electron.
2. It is circular in shape.	It has different three dimensional shapes. *Eg.* '*s*'-orbitals are spherical, *p*-orbitals are dump bell shaped etc.
3. An orbit can contain a maximum $2n^2$ electrons where n represents the order of the orbit from the nucleus.	An orbital can contain a maximum of only 2 electrons.
4. The position and momentum of the electron can be calculated at the same time.	It is not possible to find the exact position and momentum of the electron at the same time.
5. They are designated as K, L, M, N etc	They are designated as s, p, d, f etc

4.Explain spin quantum number.

Spin Quantum Number (s):

1. It represents the direction of the spin of the electrons.
2. It is denoted by the symbol s.
3. The electron may spin in the clockwise ↑ direction or anticlockwise ↓ direction. And hence it can have only two values namely either + ½ or –½.
4. Two electrons with the same sign of spin are said to have parallel spins and are represented by ↓↓ (or) ↑↑ while those having opposite spins are said to have anti parallel spins ↑↓ and are known as paired up electrons.

5.Write a note on Sommerfeld model of atom.

According to Sommerfeld, the electron moving around the nucleus must describe an elliptical orbit in addition to circular one as suggested by Bohr.

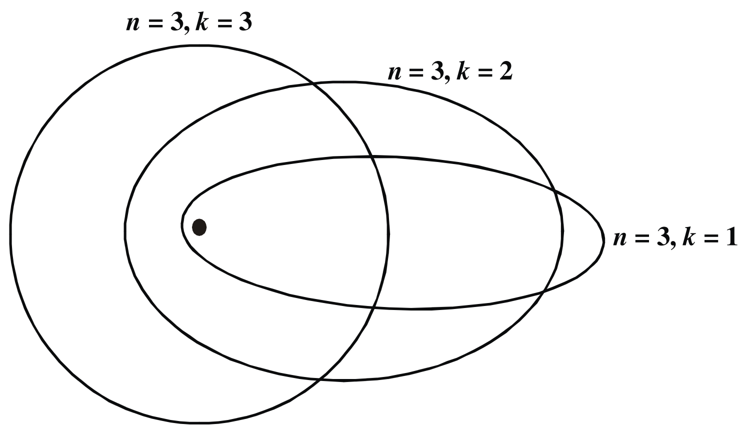

Fig .Bohr Sommerfeld orbits for n = 3, k = 1,2,3

Note: Both n and k are integers. When the value of k is equal to n, the orbit is circular, but for other values of k, the orbit is elliptical.

IV. ANSWER IN DETAIL.(T.B. Page 21)

1. **Discuss the four types of quantum numbers in detail.**

The numbers which designate and distinguish various atomic orbitals and electrons present in an atom are called quantum numbers.

In an atom, the state of each electron is different with respect to the nucleus. In order to define the state of the electron completely, four quantum numbers are used.

They are -

1. Principal quantum number (n)
2. Azimuthal quantum number (l)
3. Magnetic quantum number (m)
4. Spin quantum number (s).

1. Principal Quantum Number (*n*):

1. It determines the energy shell in which the electron is revolving around the nucleus. It is also known as major energy level.
2. It is denoted by the symbol n and may have any integral value except zero. *i.e.,* it can have the value n =1,2,3, …. etc.
3. The value n = 1 denotes that the electron is in the first shell (K shell).
   The value n = 2 denotes that the electron is in the second shell (L shell).
   The value n = 3 denotes that the electron is in the third shell (M shell).
   The value n = 4 denotes that the electron is in the fourth shell (N shell).
4. As the distance of the electron from the nucleus increases, its energy becomes higher and higher.
5. The maximum number of electrons in a major energy level is given by $2n^2$.

Principal quantum number 'n'	Designation	Maximum number of electrons ($2n^2$)
1	K	2
2	L	8
3	M	18
4	N	32

2. Azimuthal Quantum Number or Orbital Quantum Number (*l*):

1. It represents the **sub shell** to which the electron belongs.
2. It is denoted by the symbol l. Its value depends on the principal quantum number n. It may have any value ranging from 0 to (n – 1).

Principal quantum **l Value** **Name of the**

number '*n*'	*l* = (*n*-1)	sub shells or orbital
1	0	1*s*
2	0	2*s*
	1	2*p*
	0	3*s*
3	1	3*p*
	2	3*d*
	0	4*s*
4	1	4*p*
	2	4*d*
	3	4*f*

3. The value l = 0 denotes that the electron is in the s sub shell or *s* orbital.
   The value l = 1 denotes that the electron is in the p sub shell or *p* orbital.
   The value l = 2 denotes that the electron is in the d sub shell or *d* orbital.
   The value l = 3 denotes that the electron is in the f sub shell or *f* orbital.

3. Magnetic Quantum Number (*m*):

1. It represents the **orientation** of an atomic orbital in space.
2. It is denoted by the symbol *m*. The possible value which m can have depends upon the value of l. It may have all the integral values between – l to + l through 0 that is the total number of values of m would be (2l + 1).
3. Its value tells the orientations of orbital in space. The value of m = 0 denotes that the orbital has **no orientation**. The value of m = 1 denotes that it has three orbital with **three types** of orientations. The value of m = 2 denotes that it has five orbital with **five types** of orientations.

Principal quantum number '*n*'	*l* value *l* =(*n*-1)	m value (- *l* ··· O ··· + *l*)	Name of the sub shells or orbital with orientation
1	0	0	1*s*
2	0	0	2*s*
	1	– 1, 0, + 1	$2p_x$, $2p_y$, $2p_z$
	0	0	3*s*
3	1	– 1, 0, + 1	$3p_x$, $3p_y$, $3p_z$
	2	– 2, – 1, 0, +1 ,+2	$3d_{xy}$, $3d_{xz}$, $3d_{yz}$, $3d_{z^2}$, $3d_{x^2-y^2}$

4. Spin Quantum Number (*s*):

1. It represents the direction of the spin of the electrons.
2. It is denoted by the symbol s.
3. The electron may spin in the clockwise ↑ direction or anticlockwise ↓ direction. And hence it can have only two values namely either + ½ or –½.
4. Two electrons with the same sign of spin are said to have parallel spins and are represented by ↓↓ (or) ↑↑ while those having opposite spins are said to have anti parallel spins ↑↓ and are known as paired up electrons.

2.State and explain Hund's rule and Pauli's exclusion principle with suitable illustrations.

Hund's Rule: **"Among the orbitals of same energy, electrons do not start pairing, until all these orbitals are singly occupied".**

Hund's rule is also called as the principle of minimum pairing and the principle of maximum multiplicity.

Example I: In the case of nitrogen, there are 3 electrons to be filled in $2p_x$, $2p_y$ and $2p_z$ orbitals According to Hund's rule one electron will be filled in each one of these degenerate orbitals as $2p_x^1$, $2p_y^1$, $2p_z^1$.

$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑	↑	↑

Example II: In the case of oxygen, there are 4 electrons to be filled in $2p_x$, $2p_y$ and $2p_z$ orbitals. In this case the number of electrons exceeds the number of orbitals. According to Hund's rule, each one of $2p_x$, $2p_y$ and $2p_z$ is singly occupied. Afterwards, the fourth electron is filled in one of the singly occupied orbitals, but the spins of these two electrons must be opposite (↑↓). This is shown as $2p_x^2$, $2p_y^1$, $2p_z^1$.

$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑↓	↑	↑

"In an atom no two electrons can have the same set of four quantum numbers".

Illustration of Pauli's exclusion principle:

1. In an atom if one electron is assigned a set of four quantum numbers $n = 1$, $l = 0$, $m = 0$, $s = + ½$, then other electrons **cannot be assigned** the same set of quantum numbers.
2. If three quantum numbers for two electrons are the same, then these electrons must have different fourth quantum number.

	n	***l***	***m***	***s***
First electron	1	0	0	+ ½
Second electron	1	0	0	– ½

3.Discuss the shapes of *s* and *p* orbitals with a neat diagram.

***s* orbital:**

1. s orbital is spherically symmetric around the nucleus, *i.e.,* probability of finding the electron at a particular distance from the nucleus is the same in all directions.
2. 1*s* orbital does not contain any node and is the smallest of all subsequent *s* orbitals. The size of an *s* orbital increases with increase in the value of *n*.
3. 2*s* orbital is larger in size as compared to 1*s* orbital possesses a node. The 3*s* orbital still larger in size and contains two nodes. The shapes of 1*s*, 2*s* and 3*s* orbitals are shown in *Fig*.

1*s* Orbital **2*s* Orbital** **3*s* Orbital**

Fig. Shapes of 1s, 2s and 3s orbitals

***p* orbitals:**

1. *p*-orbital has three orientations *i.e.* probability of finding *p*-electron is along mutually perpendicular X, Y and Z axis. These orbitals are thus named as p_x, p_y ,and p_z orbital.
2. In p_x orbital, the electron density is distributed along X-axis while in p_y and p_z orbitals, the electron density distributions are along Y and Z axes respectively.
3. Each *p* orbital is dumb bell shaped and consists of two lobes of electron cloud which extend outwards and away from the nucleus along the axial line.
4. A nodal plane exists between the two lobes. Along this plane, the probability of finding electron (ψ^2) is zero and consequently the electron density is also zero.
5. In each *p* orbital, **the point at which the two lobes meet together is a nodal point**. It is the point from which the nodal plane passes. The shapes of $2p_x$, $2p_y$ and $2p_z$ orbitals are shown in *Fig*.

Fig. Shapes of $2p_x$, $2p_y$ and $2p_z$ orbitals

Nodes are the region in which probability of finding electrons (ψ^2) is zero.

The above four quantum numbers give the position of any electron in the major energy level, the orientation of electron in the orbital and the direction of its spin. The various states that an electron can occupy are summarized in the table given below.

OTHER IMPORTANT QUESTIONS & ANSWERS

SELF EVALUATION

I. CHOOSE THE CORRECT ANSWER:

1. In Aufbau principle, the word ‘Aufbau’
   (a) represents the name of scientist who developed the principle
   (b) is a German word which means “built up”
   (c) is related to the energy momentum of the electron
   (d) is related to angular momentum of the electron.
2. In which of the following electron distribution in ground state, only the Hund’s rule is violated?

	2s	2p		2s	2p
(a)	[↑↓]	[↑][ ][ ]	(b)	[↑]	[↑][↑][↑]
(c)	[↑↑]	[↑↓][↑][ ]	(d)	[↑↓]	[↑][↑][↑]

3. The total number of electrons in a principal energy shell is designated by expression
   (a) n (b) 2n + 1 (c) n^2 (d) $2n^2$
4. The total number of electrons in a subshell designated by azimuthal quantum number, l is given as
   (a) $2l + 1$ (b) l^2 (c) $4l + 2$ (d) $2l + 2$.

5. Wave mechanical model of the atom depends upon
   (a) de-Broglie’s concept of duality (b) Uncertainty principle
   (c) Schrodinger’s wave equation (d) All the above
6. The conclusion that orbital can accommodate only two electrons is derived from
   (a) Heisenberg’s principle (b) Aufbau rule (c) Pauli’s exclusion principle (d) Hund’s rule
7. A region in space around the nucleus of an atom where the probability of finding the electron is maximum is called
   (a) Sub-level (b) orbit (c) orbital (d) electron shell.
8. Which out of the following configurations is incorrect?
   (a) $1s^2\ 2s^2\ 2p^2_x\ 2p_y^2\ 2p^0_z$ (b) $1s^2\ 2s^2\ 2p^1_x\ 2p_y^1\ 2p^0_z$
   (c) $1s^2\ 2s^2\ 2p^1_x\ 2p_y^1\ 2p^1_z$ (d) $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6\ 3d^5\ 4s^1$
9. When 3p orbitals are completely filled then, the newly entering electron goes into
   (a) 4p (b) 3d (c) 4s (d) 4d
10. How many electrons in $_{19}$K have n = 3; l = 0?
   (a) 1 (b) 2 (c) 4 (d) 3.
11. The maximum number of 3d electrons having spin quantum s = + ½ are
   (a) 10 (b) 14 (c) 5 (d) any number from 1 to 10.
12. In which of the following all the electrons are paired?
   (a) Atom with atomic number 22 (b) nitrogen atom
   (c) Atom with configuration $3s^2\ 3p^4$ (d) magnesium
13. The valence electrons of $_{29}$Cu lie in the
   (a) K shell (b) M shell (c) N shell (d) both M and N shell.
14. The number of electrons that can be accommodated in d_{xy} orbital is
   (a) 10 (b) 4 (c) 1 (d) 2.
15. The electronic configuration of Mn can be written as
   (a) [Ar] $4s^2$ (b) [Ar] $3d^6$, $4s^2$ (c) [Ar] $3d^5$, $4s^1$ (d) [Ar] $3d^5$, $4s^2$.
16. Which of the following sets of quantum numbers is not possible for 23rd electron of Cr(at. No. (24)?
   (a) 3,2 + 2, - ½ (b) 3,2 - 2, + ½ (c) 3,2 + 1, + ½ (d) 3,1 + 1, + ½
17. How many quantum numbers are required to define the electron in atom?
   (a) two (b) three (c) one (d) four.
18. The total number of electrons present in any main energy level can be calculated from
   (a) $(2l + 1)$ (b) $2n^2$ (c) (2n + 1) (d) n^2
19. Which of the following sets of quantum numbers is allowable?
   (a) n = 2, l = 1, m = 0, s = + ½ (b) n = 2, l = 2, m = -1, s = - ½
   (c) n = 2, l = - 2, m = 1, s = + ½ (d) n = 2, l = 1, m = 0, s = 0.

20. Which shape is associated with the orbital designated by n = 2; l = 1?
(a) Spherical (b) tetrahedral (c) dumb – bell (d) pyramidal.
21. Which of the following sets of quantum numbers is impossible arrangement?
(a) n = 3,m = -2, s = + ½ (b) n = 4,m = 3, s = + ½
(c) n = 5,m = 2, s = - ½ (d) n = 3,m = -3, s = - ½ .
22. Which of the following statements about quantum numbers is wrong?
(a) If the value of l = 0, the electron distribution is spherical
(b) The shape of the orbital is given by subsidiary quantum number
(c) The Zeeman's effect is explained by magnetic quantum number
(d) The spin quantum number gives the orientations of electron cloud.
23. The shape of the orbital with the value of l =2 and m = 0 is
(a) Spherical (b) dumb- bell (c) trigonal planar (d) square planar.
24. Which of the following sets of quantum numbers is not possible?
(a) n = 4,l = 1, m = 0,s = + ½ (b) n = 4,l = 3, m = -3,s = - ½
(c) n = 4,l = 1, m = +2,s = - ½ (d) n = 4,l = 0, m = 0,s = - ½
25. The possible sub shells in n = 3 energy shell are
(a) s,p,d (b) s,p,d,f (c) s,p (d) s only.
26. If the value of m for an electron is + 3. It may be found in
(a) 4s- orbital (b) 4p-orbital (c) In any f-orbital (d) In any d- orbital
27. Which of the following orbital does not make sense?
(a) 3d (b) 3f (c) 5p (d) 7s.
28. If the value of azimuthal quantum number of an electron is 2, then which of the flowing values of magnetic quantum numbers is not permissible,
(a) 3 (b) 2 (c) 0 (d) 1
29. The quantum number which is related to the orbital angular momentum is
(a) subsidiary quantum number (b) principle quantum number
(c) magnetic quantum number (d) spin quantum number.
30. The value of azimuthal quantum number for electrons present in 6p-orbital is
(a) 2 (b) 1 (c) any of the value between 0 and 5 (d) 0.
31. Which of the following is the correct set of quantum numbers for the outer shell electrons of $_{21}$Sc?
(a) 3,2,0, + ½ (b) 4,0,0, + ½ (c) 3,0,0, - ½ (d) 4,0, -1, + ½ .
32. Which value of l will represent double dumb- bell shape of the orbital?
(a) 0 (b) 1 (c) 2 (d) l does not give shape of orbital.
33. How man electrons in K (Z = 19) have n= 4;l = 0?
(a) 1 (b) 2 (c) 3 (d) 4.
34. Indicate which electronic configuration amongst the following correctly represent SULPHUR atom?
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3d^2$ (b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2\ 4s^2$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 4p^1$ (d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
35. The magnetic quantum number represents
(a) Size of the orbital (b) spin angular momentum
(c) orbital angular momentum (d) spatial orientation of orbital.
36. No two electrons in an atom will have all the four quantum numbers same. This statement is know as
(a) Exclusion principle (b) Uncertainty principle (c) Hund's rule (d) Aufbau principle.
37. The maximum number of electrons in a subshell for which l = 3 is
(a) 14 (b) 10 (c) 8 (d) 4.
38. The number of electrons in the M shell of the element with atomic number 24 is
(a) 24 (b) 12 (c) 13 (d) 8.
39. The two electrons occupying an orbital are distinguished by
(a) principal quantum number (b) azimuthal quantum number
(c) magnetic quantum number (d) spin quantum number.
40. The symbol of the element whose atoms have the outer most electronic configuration $2s^2\ 2p^3$ is
(a) N (b) Li (c) P (d) Na.
41. The principal quantum number, n describes
(a) Shape of orbital (b) sub- shell of electron
(c) Main energy shell of electron (d) spin of electron.
42. The quantum numbers for the outer electrons of an atom are given by
n = 2; l = 0; m = 0 ; s = + ½ is given by
(a) Lithium (b) Beryllium (c) Hydrogen (d) Boron.
43. $^{22}_{11}$Na contains
(a) 22 protons (b) 11 neutrons (c) 22 neutrons (d) None of these.
44. Which quantum number is sufficient to describe the electron in hydrogen atom?

(a) l (b) n (c) m (d) s.

45. Which one of the following shows the correct electronic configuration of the outermost shell in innert gases?
(a) ns^2, np^6 (b) ns^2, np^3 (c) ns^2,np^5 (d) ns^2, np^4.

46. The valence orbital configuration of an element with atomic number 23 is
(a) $3d^5$ (b) $3d^3$, $4s^2$ (c) $3d^3$, $4s^1$, $4p^1$ (d) $3d^2$, $4s^2$, $4p^1$.

47. The electronic configuration of $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^2$ corresponds to
(a) Si (b) S (c) Na (d) Ar.

48. When the azimuthal quantum number l =1, the shape of the orbital will be
(a) Spherical (b) dumb – bell (c) double dumb – bell (d) more complicated.

49. The correct set of quantum numbers for the unpaired electron of chlorine atom is
(a) n = 2, l = 1, m =0 (b) n = 2, l = 1,m =1 (c) n = 3, l = 1,m =1 (d) n =3, l = 0,m = 0.

50. Oxidation involves:
(a) gain of electrons (b) loss of electrons
(c) increase in the valency of negative part (d) decrease in the valency of positive part

51. Reduction involves:
(a) gain of electrons (b) loss of electrons
(c) increase in the valency of negative part (d) decrease in the valency of positive part

52. The oxidation number of Cr in $K_2Cr_2O_7$:
(a) +2 (b) –2 (c) +6 (d) –6

53. The oxidation number of carbon in CH_2O is:
(a) –2 (b) +2 (c) 0 (d) +4

54. The oxidation number of Mn in MnO_4^- is
(a) +7 (b) –5 (c) –7 (d) +5

55. The oxidation number of carbon in $CHCl_3$ is:
(a) +2 (b) +3 (c) +4 (d) –3

56. The oxidation state of sulphur in SO_4^{2-} is:
(a) +2 (b) +4 (c) +6 (d) –6

57. The oxidation state of sulphur in S_2O7^{2-} is:
(a) +6 (b) –6 (c) –2 (d) +2

58. Oxidation number of oxygen in H_2O and H_2O_2 are respectively:
(a) +2 and –1 (b) –2 and +2 (c) –2 and –2 (d) –2 and -1

59. Oxidation number of sulphur in H_2SO_4 is:
(a) +2 (b) +4 (c) +6 (d) +8

60. Which of the following sets of the quantum numbers is permitted?
(a) n = 4, l = 2, m = +3, s = + 1/2 (b) n = 3, l = 3, m = +3 , s = +1/2
(c) n = 4, l = 0, m = 0, s = + 1/2 (d) n = 4, l = 3, m = +1 , s = 0

61. A sub shell with n = 6, l = 2 can accommodate a maximum of
(a) 12 electrons (b) 36 electrons (c) 10 electrons (d) 72 electrons

62. An electron has spin quantum number, s = + 1/2 and magnetic quantum number, m = + 1. It cannot be present in
(a) s- orbital (b) p-orbital (c) d-orbital (d) f-orbital

63. According to Aufbau principle, the 19^{th} electron in an atom goes into the
(a) 4s- orbital (b) 3d-orbital (c) 4p-orbital (d) 3p-orbital

64. How many electrons in calcium have l = 0?
(a) 6 (b) 8 (c) 10 (d) 12

65. The number of electrons with quantum numbers n = 3 and l = 2 in chromium is
(a) 8 (b) 6 (c) 5 (d) 7

66. Neutrons were discovered by
(a) J.J Thomson (b) Rutherford (c) James chadwick (d) G.T. Seabery

67. The orbital diagram in which the Aufbau principle is violated

	2s	2p				2s	2p		
(a)	↑	↑↓	↑		(b)	↑	↑↓	↑	↑
(c)	↓	↑	↑	↑	(d)	↑	↑↓	↑↓	↑

68. A p-orbital can accommodate
(a) 4 electrons (b) 6 electrons
(c) 2 electrons with parallel spin (d) 2 electrons with opposite spin

69. The principal quantum number of an atom represents

(a) Size of orbital (b) spin angular momentum
(c) Orbital angular momentum (d) space orientation of the orbital

70. The maximum number of electrons in an atom which can have n = 4 is

(a) 4 (b) 8 (c) 16 (d) 32

71. The maximum value of l for an electron in fifth energy level is

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

72. If the value of principal quantum number is 3, the total possible values for magnetic quantum number will be

(a) 1 (b) 4 (c) 9 (d) 12

73. Azimuthal quantum number of last electron of $_{11}Na$ is

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

74. The correct set of quantum numbers for the unpaired electron of chlorine atom is

	n	l	m
(a)	2	1	0
(b)	2	1	1
(c)	3	1	1
(d)	3	0	1

5. Which quantum number will determine the shape of the sub shell?

(a) Principal quantum number (b) azimuthal quantum number
(c) Magnetic quantum number (d) Spin quantum number

76. The four quantum numbers of valence electron of potassium are

(a) 4,0,1, 1/2 (b) 4,1,0,1/2 (c) 4,0,0, 1/2 (d) 4,1,1,1/2

77. An electron is present in 4f sub-shell. The possible values of azimuthal quantum number for this electron are

(a) 0,1,2,3 (b) 1,2,3,4 (c) 3 (d) 4

78. Which of the following electronic configuration is correct?

(a) $1s^2, 2s^2, 2p_x^2, 2p_y^2, 2p_z^2, 3s^2, 3p_x^2$ (b) $1s^2, 2s^1, 2p_x^1, 2p_y^1, 2p_z^1$
(c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^4, 4s^2$ (d) $1s^2, 2s^2, 2p^6, 3s^2, 3p_x^1, 3p_y^1, 3p_z^1$

79. Which of the following is not true for a principal energy level having n = 3?

(a) There are three sub-shells (b) There are nine orbital
(c) There are a maximum of 18 electrons (d) There are six electrons with l = 2

80. The maximum number of electrons in p-orbital with n = 6 m = 0 is

(a) 2 (b) 6 (c) 10 (d) 14

81. Which of the following sets of quantum numbers is correct?

(a) n = 4, l = 3, m = + 4, x = +1/2 (b) n = 3, l = 2, m = + 3, s = -1/2
(c) n = 2, l = 2, m = + 2, s = +1/2 (d) n = 1, l = 0, m = 0, s = -1/2

82. For n = 4

(a) the total possible values of l are 3 (b) the highest value of l is 4
(c) the total number of possible values of m is 7 (d) the highest value of m is + 3

83. Which of the following orbital designation is not possible

(a) 5*f* (b) 5g (c) 7*p* (d) 3*f*

84. Nitrogen has the electronic configuration $1s^2\ 2s^2\ 2p^1_x\ 2p^1_y\ 2p^1_z$ and not $1s^2\ 2s^2\ 2p^2_x\ 2p^1_y\ 2p^0_z$. This is determined by

(a) Pauli's exclusion principle (b) Aufbau principle
(c) uncertainty principle (d) Hund's rule

85. The set of quantum numbers *n*, *1* and *m* for the valence electron of sodium (atomic number 11) is

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

86. The azimuthal quantum number value for 3p electron is

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

87. The fact that an orbital can accommodate a maximum of two electrons is deduced from

(a) Hund's rule (b) Pauli's exclusion principle
(c) Aufbau principle (d) uncertainty principle

88. Which of the following value of *1* is not possible for n = 4?

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

89. The number of orbitals in n = 3 quantum level is

(a) 6 (b) 18 (c) 9 (d) 12

90. The set of quantum numbers not applicable to an electron in an atom is

(a) n = 1, l = 0, m = 0, s = + ½ (b) n = 1, l = 1, m = 0, s = + ½
(c) n = 1, l = 0, m = 0, s = - ½ (d) n = 2, l = 1, m = 1, s = + ½

91. The energy of an electron is mainly determined by

(a) principal quantum number (b) azimuthal quantum number

(c) magnetic quantum number (d) spin quantum number

92. The number of orbitals possible in a sub-shell with $l = 3$ is
(a) 3 (b) 5 (c) 6 (d) 7

93. Maximum number of electrons that can be placed in 2p sub-shell is
(a) 2 (b) 4 (c) 6 (d) 8

94. Among the following which is the correct outermost configuration of Cu atom?
(a) $4s^2\ 3d^9$ (b) $4s^1\ 3d^{10}$ (c) $4s^1\ 3d^9$ (d) $4s^1\ 3d^8$

95. The oxidation number of S in S^{-2} is:
(a) –2 (b) 0 (c) –6 (d) +2

96. Magnetic quantum number specifies
(a) size of orbitals (b) shape of orbitals
(c) orientation of orbitals in space (d) spin angular momentum

97. Magnetic quantum number for the last electron in potassium is
(a) 0 (b) 1 (c) 2 (d) 3

98. Oxidation may be defined as:
(a) Addition of electron (b) Gain of electron
(c) Addition of hydrogen (d) Addition of oxygen

99. In an oxidation process, oxidation number:
(a) Decreases (b) Increases
(c) Does not change (d) First increases then decreases

100. The process in which oxidation number increases is known as:
(a) Oxidation (b) Reduction (c) Auto oxidation (d) None of the above

101. The oxidation number of hydrogen in LiH is:
(a) +1 (b) –1 (c) 2 (d) 0

102. The oxidation number of oxygen in O_2 molecule is:
(a) 0 (b) –1/2 (c) –2 (d) +2

103. Oxidation State of oxygen in hydrogen peroxide is:
(a) –1 (b) +1 (c) 0 (d) –2

104. The Oxidation State of chlorine in $KClO_4$ is:
(a) –1 (b) +1 (c) +7 (d) –7

105. The Oxidation State of Mn in K_2MnO_4 is:
(a) +6 (b) –6 (c) +2 (d) –2

106. The oxidation number of N in NH_3 is:
(a) +3 (b) –3 (c) +5 (d) 0

Answers:

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

I. ANSWER IN ONE OR TWO SENTENCES:

1. What is meant by an anu?

If an element is broken down into smaller and smaller and the process would end at the smallest particle which cannot be broken any further and this particle is called an anu.

2. What is an atom?

An atom is an extremely small and indivisible particle. All matters are composed of such very tiny particles called atoms. Atom is a Greek word that means indivisible.

3. Why Principal quantum number is called as major energy level?

Principal quantum number determines the energy shell in which the electron is revolving around the nucleus. Hence it is also known as major energy level.

4. What is principal quantum number?

Principal quantum number determines the energy shell in which the electron is revolving around the nucleus. It is represented by the symbol ***n*** and may have any integral value except zero.

5. What is Azimuthal quantum number?

Azimuthal quantum number represents the sub shell to which the electron belongs. It is denoted by the symbol 'l' and may have any value form 0 to (n-1).

6. What is magnetic quantum number?

It represents the orientation of an atomic orbital in space and is denoted by the symbol 'm'. 'm' depends upon the value of 'l'. It may have all the integral values between-1 to +1 through 0 that is the total number of values 'm' would be $(2l + 1)$.

7. What is the orientation when m = 0?

The value of $m = 0$ denoted that the orbital has no orientation.

8. How many orientations are possible when m = 1?

The value of $m = 1$ denotes that it has three orbital with three types of orientations. They are p_x, p_y and p_z. (i e l = - 1,0,+1)

9. Write the orientations when the value of m = 2?

The value of m = 2 denotes that it has five orbitals with five types of orientations. They are d_{xy}, d_{yz}, d_{zx}, d_{x2-y2} and d_{z2}. (i e l = -2,-1,0,+1,+2)

10. What is spin quantum number?

Spin quantum number represents the direction of the spin of the electrons. The electron may spin in the clockwise↑ direction or anticlockwise ↓ direction. Hence it can have two values namely either +1/2 or -1/2.

11. Draw the structure of an orbital and or orbit.

Diagram

12. Explain the term dual character of electron?

An electron behaves as if it is a particle as well as a wave. This is known as the dual nature of the electron

13. How many orientations are possible for a *p* sub shell?

The 'p' subshell has three possible orientations. One along the x axis, (p_x) the other along the y axis (b_y) and the other along the z axis.(p_z)

14. A '2*p*' orbital is filled first and then '3*s*' orbital why? Give reason?

For 2p orbital $n + l = 2 + 1 = 3$.

For 3s orbital $n + l = 3 + 0 = 3$.

According to (n + l) rule, '2p' orbital is filled first before the '3s' orbital.

15. Among 3*d* and 4*p*, which orbital is filled first? justfy your answer.

For a 3d orbital: $n + l = 3 + 2 = 5$.

For a 4s orbital: $n + l = 4 + 1 = 5$.

According to ($n + l$) rule, the '3d' orbitals are filled first.

16. Arrange the following orbitals in their increasing order of their (*n* + *l*) value and hence their energies? 4*d*, 5*p*, 4*f*.

($n + l$) value for 4d orbital = 4 + 2 = 6.

($n + l$) value for 5p orbital = 5 + 1 = 6.

($n + l$) value for 4f orbital = 4 + 3 = 7.

Thus 4f orbital has the highest energy, among 4d and 5p orbitals, 4d orbital has lesser energy. Hence the increasing order of energy is - 4d< 5p< 4f.

18. Using the box diagram, show the electrons in the appropriate orbitals, for Berylium (Z = 4) and Fluorine (Z =9)?

Berylium $1s^2$ $2s^2$

↑↓	↑↓

Fluorine	$1s^2$	$2s^2$	$2p^5$		
	↑↓	↑↓	↑↓	↑↓	↑

19. Identify the oxidising agent and reducing agent in the following reactions?

(i) $H_2S + CI_2 \longrightarrow 2HCI + S$

increase in ON

$$H_2S + C1_2 \longrightarrow 2HC1 + S$$

Decrease in ON

The S^{-2} in H_2S is oxidised to S. Hence it is a reducing agent. The oxidation number of chlorine is decreased from zero to – 1. Hence it is reduced to C1-. It is the oxidising agent.

20. What is the oxidation number of Mn in $KMnO_4$?

Oxidation number of K is +1 and O is -2. So,

$+1 + x - 8 = 0$

$x - 7 = 0$

$x = +7$

The oxidation number of Mn in $KMnO_4$ = +7.

21. What are energy levels?

The electrons is an atom revolve around the nucleus in two dimensional circular paths called orbits which have a definite energy. These orbits with fixed energies are called the energy levels. They are K, L, M, N, etc.

22. What are sub shells?

A given energy level consists of a definite number of sub levels with slightly different energies. These sub levels of the given energy level are called sub-shells. They are *s, p, d, f…*

23. Give the electronic configuration of aluminium?

$1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^1$

24. Write the electronic configuration of an element with atomic number 16^s?

$1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^4$

25. Calculate the oxidation number of Cr in $Cr_2O_7^{2-}$?

Oxidation number of oxygen is -2

$2x + (7x - 2) = -2$

$2x - 4 = -2$

$2x = 12$

$x = +6$

The oxidation number of Cr in$Cr_2O_7^{2-}$ is +7

26. An atomic orbital has (i) n = 4,(ii) n=2 What are the possible values of l?

(i) n = 4, the possible values of l will be between 0 and 3 Therefore the possible values of l are 0, 1, 2, 3.
(ii) n = 2, the possible values of l will be between 0 and 1. Therefore the possible values of l are 0, 1

27. Name the various types of atomic models.

1. Daltons model (1808)
2. J.J. Thomson model (1897)
3. Rutherford's model (1911)

4. Bohr's model (1913)
5. Sommerfeld model
6. Quantum mechanical model of atom

28. Define orbit.

An orbit is a definite circular path, whereas an orbital represents the total volume where the electron spends most of its time.

29. Define orbital.

An orbital may be defined as a region in the three dimensional space around the nucleus where the probability of finding the electron is maximum.

30. Name the three fundamental particles of an atom? Who found this?

In 1886 Goldstein discovered proton (${}_{1}H^{1}$).

In 1897 J.J. Thomson discovered electron (${}_{-1}e^{0}$)

In 1932 Chadwick discovered neutron (${}_{0}n^{1}$).

31. What are quantum numbers?

The numbers which designate and distinguish various atomic orbitals and electrons present in an atom are called quantum numbers.

32. What are the four types of quantum numbers?

(i) Principal quantum number (n)
(ii) Azimuthal quantum number (l)
(iii) Magnetic quantum number (m)
(iv) Spin quantum number (s).

33. Define electronic configuration.

Distribution of electrons in different orbitals of the atom of an element is called **electronic configuration**.

34. What do you mean by redox reaction? Give example.

Chemical reactions involving simultaneous oxidation and reduction are called 'redox reactions'.

Oxidation

Example: $Fe^{3+} + Sn^{2+} \longrightarrow Fe^{2+} + Sn^{4+}$

Reduction

35. Define Reducing agent. Give example.

A substance which gives one or more electrons to the other is called a reducing agent. In the above example, Na is the **reducing agent**.

$$Na \longrightarrow Na^{+} + e^{-}$$

36. Define oxidizing agent. Give example.

A substance which accepts one or more electrons from the other is called an oxidizing agent.

$$F + e^{-} \longrightarrow F^{-}$$

In the above example, Cl is the **oxidizing agent**.

37. What are nodal points and planes?

An orbital contain one or more points or planes where the probability of finding electron is zero. At such points or planes, electron density is found to be **equal to zero**. Such points are called **nodal points** and such planes are termed as **nodal planes**. The shapes of some orbitals are discussed below.

III. ANSWER IN BRIEF:

1. Discuss the importance at Hund's rule in detail.

Importance of Hund's rule:

This rule provides guidelines of filling electrons in the degenerate orbitals (orbitals having equal energies) of an atom. According to this rule:

- If the number of electrons is equal to (or less than) the number of degenerate orbitals, then orbitals are singly occupied.
- Pairing of electron spins takes place only when each one of the degenerate orbitals is singly occupied. This is possible only when the number of electrons **to be filled is greater than the number of degenerate orbitals.**

Example I: In the case of nitrogen, there are 3 electrons to be filled in $2p_x$, $2p_y$ and $2p_z$ orbitals. According to Hund's rule one electron will be filled in each one of these degenerate orbitals as $2p_x^1$, $2p_y^1$, $2p_z^1$.

$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑	↑	↑

Example II: In the case of oxygen, there are 4 electrons to be filled in $2p_x$, $2p_y$ and $2p_z$ orbitals. In this case the number of electrons exceeds the number of orbitals. According to Hund's rule, each one of $2p_x$, $2p_y$ and $2p_z$ is singly occupied. Afterwards, the fourth electron is filled in one of the singly occupied orbitals, but the spins of these two electrons must be opposite (↑↓). This is shown as $2p_x^2$, $2p_y^1$, $2p_z^1$.

$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑↓	↑	↑

2. Give the electronic configuration of element whose atomic numbers from 11 to 20.

Similarly the electronic configuration of atom from Sodium to Zinc is as given below.

Name	**Element Symbol**	**Z Atomic Number**	**Electronic Configuration**
Sodium	Na	11	[Ne] $3s^1$
Magnesium	Mg	12	[Ne] $3s^2$
Aluminium	Al	13	[Ne] $3s^2$ $3p^1$
Silicon	Si	14	[Ne] $3s^2$ $3p^2$
Phosphorus	P	15	[Ne] $3s^2$ $3p^3$
Sulphur	S	16	[Ne] $3s^2$ $3p^4$
Chlorine	Cl	17	[Ne] $3s^2$ $3p^5$
Argon	**Ar**	**18**	**$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ ≡ [Ar]**
Pottassium	K	19	[Ar] $4s^1$
Calcium	Ca	20	[Ar] $4s^2$

3. Explain oxidation and reduction in terms of oxidation number.

A process in which the oxidation number of an element increases is called **oxidation.**

A process in which the oxidation number of an element decreases is called **reduction**.

eg: In the reaction,

$$2Fe^{+3} + Sn^{+2} \longrightarrow 2Fe^{+2} + Sn^{+4}$$

Oxidation number of iron decreases from + 3 to + 2. It is reduction.

Oxidation number of tin increases from +2 to +4. It is oxidation.

This is a redox reaction in which ferric ion is reduced into ferrous ion and tin is oxidized from stannous form to stannic.

4. What is wave mechanical model of an atom.

Based on the dual character of an electron and Heisenberg's Uncertainty Principle. E. Schrodinger in 1926 formulated a new model called quantum mechanical model of atom or wave mechanical model of atom. He considered the electron as **three-dimensional wave** in electric field of the positively charged nucleus. To describe the behaviour of electron waves, Schrodinger developed a mathematical equation known as Schrodinger wave equation.

IV. ANSWER IN DETAIL:

1. Define oxidation number. Discuss the rules of oxidation number.

Oxidation number: Oxidation number is defined as the actual or apparent charge possessed by an atom of the element in a compound.

Various rules to determine oxidation number are given below:

(i) In the elementary state, oxidation number of an atom is zero.
eg. O_2, H_2, S, Na etc.

(ii) In a monoatomic ion, charge on the ion itself represents oxidation number.

eg. Oxidation numbers of Na^+, Ba^{2+}, Al^{3+}, Cl^-, S^{2-} are +1, +2, +3, –1, –2 respectively.

(iii) Fluorine, the most electronegative element has an oxidation number –1, in all its compounds.

(iv) Oxidation number of hydrogen is always +1, in its compounds except metal hydrides (*eg.* LiH, NaH) in which it is –1.

(v) Oxidation number of oxygen is –2 in all its compounds except in peroxides (*eg.* H_2O_2, Na_2O_2) and in OF_2 where it is –1 and +2 respectively.

(vi) Oxidation number of alkali metals (Li, Na, K, Rb, Cs and Fr) is always +1 in their compounds.

(vii) Oxidation number of alkaline earth metals (Be, Mg, Ca, Ba, Sr, Ra) is always +2 in their compounds.

(viii) In a molecule, the sum of oxidation numbers of all the atoms should be zero.

(ix) In an ion, the sum of oxidation numbers of all the atoms is equal to charge on the ion.

2.Tabulate the various l, m values and the total number of electrons accommodated in K, L, M, N shells.

Principal Quantum Number n	Azimuthal Quantum Number l	Magnetic Quantum Number m	Total Number of electrons $2n^2$
	$l = 0$; $(1s)$	$m = 0$	$2 \times 1^2 = 2$
	$l = 0$; $(2s)$	$m = 0$	
	$l = 1$; $(2p)$	$m = -1, 0, +1$	$2 \times 2^2 = 8$
	$l = 0$; $(3s)$	$m = 0$	
	$l = 1$; $(3p)$	$m = -1, 0, +1$	$2 \times 3^2 = 18$
	$l = 2$; $(3d)$	$m = -2, -1, 0, +1, +2$	
	$l = 0$; $(4s)$	$m = 0$	
	$l = 1$; $(4p)$	$m = -1, 0, +1$	$2 \times 4^2 = 32$
	$l = 2$; $(4d)$	$m = -2, -1, 0, +1, +2$	
	$l = 3$; $(4f)$	$m = -3, -2, -1, 0, +1, +2, +3$	

TEXTBOOK PROBLEMS

NUMERICAL PROBLEMS IN QUANTAM NUMBERS (T.B. PAGE 15 & 16)

Example 1:

(a) An atomic orbital has $n = 3$. What are the possible values of l?

(b) An atomic orbital has $l = 3$. What are the possible values of m?

(c) An atomic orbital has $n = 2$. What are the possible values of l and m?

Solution:

(a) When $n = 3$, the possible values of l will be between 0 and 2. Therefore, the possible values of l will be 0, 1 and 2.

(b) When $l = 3$, the possible values of m are –3, – 2, – 1, 0, + 1, +2 and + 3.

(c) When n =2, the possible values of l are 0 and 1 while those of m are – 1, 0, +1.

Example 2:

Using the *s, p, d, f* notations, describe the orbitals with the following quantum numbers:

(a) $n = 1, l = 0$ *(b)* $n = 2, l = 1$ *(c)* $n = 3, l = 0$

(d) $n = 4, l = 3$ *(e)* $n = 5, l = 2$

Solution:

(a) $1s$ *(b)* $2p$ *(c)* $3s$

(d) $4f$ *(e)* $5d$

Example 3:

Designate the electrons having following sets of quantum numbers:

(a) $n = 2, l = 0, m = 0, s = +½$

(b) $n = 3, l = 1, m = +1, s = +½$

(c) $n = 1, l = 0, m = 0, s = -½$

Solution:

(a) The electron is present in $2s$ orbital and possesses a clockwise spin.

(b) The electron is present in $3p_z$ orbital with a clockwise spin.

(c) The electron is present in $1s$ orbital and possesses an anticlockwise spin.

Example 4:

Mention the values of n and l corresponding to the following orbitals.

(a) $1s$ *(b)* $2p$ *(c)* $3d$

(d) $4f$ *(e)* $3s$

Solution:

(a) $n = 1, l = 0$ *(c)* $n = 3, l = 2$

(b) $n = 2, l = 1$ *(d)* $n = 4, 1 = 3$

(e) $n = 3, l = 0$

(TEXT BOOK PAGE 18 & 19)

Problem 1:

Calculate the oxidation number of sulphur in the following cases:

(a) H_2S **(b) $H_2S_2O_7$**

Solution:

Let the oxidation number of S be x.

(a) **H_2S:** Oxidation number of H is +1.

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

(b) **$H_2S_2O_7$:** Oxidation number of H is +1 and O is –2.

$$\therefore (+1 \times 2) + 2x + (-2 \times 7) = 0$$
$$x = +6$$

Problem 2:

Calculate the oxidation number of chlorine in the following cases:

(a) $NaClO_3$ **(b) ClO_4^-**

Solution:

Let the oxidation number of Cl be x.

(a) **$NaClO_3$ (sodium chlorate):** Oxidation number of Na is +1 and O is –2.

$$\therefore (+1) + x + (-2 \times 3) = 0$$
$$x = +5$$

(b) ClO^-_4 (perchlorate ion): Oxidation number of O is –2.

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

ADDITIONAL PROBLEMS

1.Calculate the oxidation number of:

(a) Mn in $KmnO_4$ **(b) Mn in MnO_2**

(c) O in H_2O **(d) S in H_2SO_4**

Solution:

(a) Oxidation number of Mn in KmnO4 oxidation number of K is +1 and oxygen is -2

$$1\,(+1) + x + 4\,(-2) = 0$$
$$1 + x - 8 = 0$$
$$x - 7 = 0$$
$$x = +7$$

$\therefore$ The oxidation number of Mn in $KmnO_4$ is +7.

(b) Oxidation number of Mn in MnO_2

$$\text{Oxidation number of oxygen is} = -2$$
$$\text{Oxidation no of Mn} = x$$
$$x + 2\,(-2) = 0$$
$$x - 4 = 0$$
$$x = +4$$

$\therefore$ Oxidation no of Mn in MnO_2 is +4.

(c) Oxidation number of O in H_2O oxidation no of Hydrogen is +1

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

$$x = -2$$

$\therefore$ Oxidation number of 'O' in H_2O is –2.

(d) Oxidation number of 'S' in H_2SO_4 oxidation number of Hydrogen is +1 and oxygen is –2.

$$\begin{aligned} 2\,(H) + 1(S) + 4(O) &= 0 \\ 2\,(+1) + 1(x) + 4\,(-2) &= 0 \\ 2 + x &= 8 = 0 \\ x - 6 &= 0 \\ x &= +6 \end{aligned}$$

$\therefore$ Oxidation no of 'S' in H_2SO_4 is equal to '+6'.

2. Calculate the oxidation no of:

(a) Na in Na_2O_2 **(b) A*l* in Al_2O_3**

(c) N in HNO_3 **(d) C*l* in HOC*l*** **(e) C in Co_2**

Solution:

(a) Oxidation number of Na in Na_2O_2

Oxidation number of CO2 is – 2

$$\begin{aligned} 2\,(x) + 2\,(-2) &= 0 \\ 2x - 4 &= 0 \\ 2x &= 4 \\ x &= \\ \therefore x &= +2. \end{aligned}$$

$\therefore$ Oxidation number of Na in Na_2O_2 is = +2.

(b) Oxidation number of A*l* in Al_2O_3 Oxidation umber of O is –2

$$\begin{aligned} 2\,(x)\ + 3\,(-2) &= 0 \\ 2\,x - 6 &= 0 \\ 2x &= 6 \\ x &= \\ x &= +3 \end{aligned}$$

$\therefore$ Oxidation no of Al in Al_2O_3 is +3.

(c) Oxidation number of N in HNO_3

Oxidation number of Hydrogen (H) is +1 and oxygen (O) is –2.

$$\begin{aligned} 1\,(1) + 1\,(x) + 3(-2) &= 0 \\ 1 + x - 6 &= 0 \\ x - 5 &= 0 \\ x &= +5 \end{aligned}$$

$\therefore$ Oxidation number of N in HNO_3 is +5.

(d) Oxidation number of C*l* in HOC*l*.

Oxidation number of H is +1 and O is –2

$$\begin{aligned} 1\,(+1) + 1\,(-2) + x &= 0 \\ 1 - 2 + x &= 0 \\ x - 1 &= 0 \\ x &= +1 \end{aligned}$$

$\therefore$ Oxidation number of X in HOC*l* is +1.

(e) Oxidation number of C in CO_2

Oxidation number of O is –2

$$\begin{aligned} \therefore\ 1\,(x) + 2\,(-2) &= 0 \\ x - 4 &= 0 \\ x &= +4 \end{aligned}$$

$\therefore$ Oxidation number of 'C' in CO_2 is +4.

3. Write down the electronic configurations of atoms having atomic number 7, 11, 15 and 20.

Solution: The electronic configurations of atoms are:

At. No.= 7: $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ At. No.= 11: $1s^2\ 2s^2\ 2p^6\ 3s^1$

At. No.= 15: $1s^2\ 2s^2\ 2p^6;\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$ At. No.= 20: $1s^2\ 2p^2\ 2p^6;\ 3s^2\ 3p^6\ 4s^2$

4. Which of the following do and which do not make sense?

7p, 2d, $3s^3$, $3p_y^3$, 4f.

Solution:

7p :makes sense (There are no d-orbitals for $n = 2$)

2d : does not make sense

$3s^3$: does not make sense (3s-orbital cannot have more than two electrons)

$3p_y^3$: does not make sense ($3p_y$ orbital cannot have more than two electrons)

4f : makes sense

5. An atom has 2 electrons in the first (K) shell, 8 electrons in the second (L) shell and 2 electrons in the third (M) shell. Give its electronic configuration and find out the following:

(a) Atomic number (b) Total number of principal quantum numbers (c) Total number of sub-levels (d) Total number of s –orbitals (e) Total number of p-electrons.

Solution: The electronic configuration of the atom is:

$1s^2\ 2s^2\ 2p^6\ 3s^2$

(a) Atomic number = 2 + 2 + 6 + 2 = 12

(b) Number of principal quantum numbers = 3

(c) Number of sub –levels = 4 (1s, 2s, 2p, 3s)

(d) Number of s-orbitals = 3 (1s, 2s, 3s)

(e) Total number of p-electrons = 6.

6. What is the maximum number of unpaired electrons in Cu (Z = 29), Br^- (Z = 35) and K^+ (Z = 19)?
Solution:

Cu (Z = 29)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$	**Unpaired electrons = 1**
Br (Z = 35)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$	
Br^- (36e)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$	**Unpaired electrons = 0**
K (Z = 19)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	
K^+ (18e)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	**Unpaired electrons = 0**

7. Arrange the electrons represented by the following sets of quantum numbers in the decreasing order of energy for a multielectron atom:

(i)n = 4, l = 0, m = 0, s = + ½ (ii) n = 3, l = 1, m = 1, s = - ½
(iii) n = 3, l = 2, m = 0, s = + ½ (iv) n = 3, l = 0, m = 0, s = - ½

Solution. These represent the subshells

(i) 4s (ii) 3p (iii) 3d (iv) 3s

Order of decreasing energy:

3d > 4s > 3p > 3s or (iii) > (i) > (ii) > (iv)

8. From the following sets of quantum numbers, state which are possible. Explain why the others are not possible:

(i) n = 0, l = 0, m = 0, s = + ½ (ii) n = 1, l = 0, m = 0, s = - ½ (iii) n = 1, l = 1, m = 0, s = + ½
(iv) n = 1, l = 0, m = +1, s = + ½ (v) n = 2, l = 1, m = -1, s = - ½ (vi) n = 2, l = 2, m = 0, s = - ½

Solution: From the knowledge of possible values of quantum numbers, we can predict the possibilities of various sets of quantum numbers (ii), (v), (vi) are possible. The others are not possible because of the following reasons:

(i) n = 0, l = 0, m = 0, s = + ½	:	n cannot have value equal to zero
(iii) n = 1, l = 1, m = 0, s = + ½	:	For n = 1, l cannot be equal to 1
(iv) n = 1, l = 0, m = +1, s = + ½	:	For l = 0, m cannot be equal to +1
(vi) n = 2, l = 2, m = 0, s = - ½	:	For n = 2, l cannot be equal to 2.

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2. ATOMIC, MOLECULAR AND EQUIVALENT MASSES

INTRODUCTION:

EQUIVALENT WEIGHT

Since hydrogen is the lightest of all elements, it was chosen as a standard for determination of equivalent weights. On this basis, **the weight of an element that combines with one part by weight of hydrogen was called as its equivalent weight**. For example, in water, 16 parts by weight of oxygen combine with 2 parts by weight of hydrogen. Therefore 8 parts by weight of oxygen will combine with 1 part by weight of hydrogen. Therefore equivalent weight of oxygen is 8. Similarly, in hydrogen chloride, 35.5 parts by weight of chlorine combine with 1 part by weight of hydrogen. Therefore the equivalent weight of chlorine is 35.5.

However, hydrogen combines directly with only a few elements. On the other hand, oxygen combines directly with most of the elements. Therefore, oxygen was chosen as a standard for assigning equivalent weights to elements.

MOLE CONCEPT

A chemist needed a unit that contains a fixed number of particles of any substance which could be readily weighed. This not only will help in comparing the weights of different substance but will also facilitate the stoichiometric calculations. **This unit is called a *MOLE*.** It consists of 6.023×10^{23} particles. Counting of such a large number of particles is impossible. Therefore the concept of mole is based on a chosen weight rather than a chosen number of particles.

IMPORTANT TERMS AND DEFINITIONS

Avogadro number (N): It is defined as the number of atoms present in exactly 12 grams of ${}_{6}C^{12}$ isotope. It is denoted by N. It has a value of 6.023×10^{23}.

Atomicity. The number of atoms contained in one molecule of the element is called its atomicity.

Avogadro's hypothesis. Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

Vapour density. Vapour Density is defined as the ratio of the mass of a certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure.

Atomic weight. The relative atomic mass of an element is the mass of one atom of the element compared with the mass of one atom of hydrogen taken as one unit.

Gram atomic weight of an atom. The atomic weight of an element expressed in grams is known as the gram atomic weight (or gram atom) of the element.

For example, Gram atomic weight of carbon = 12 g

Gram atomic weight of oxygen = 16 g.

Molecular mass. The relative molecular mass of an element or a compound is the mass of one molecule of the element or compound compared with the mass of one atom of hydrogen taken as one unit.

Mole concept.

Definition 1: The mole is the amount of substance, which contains as many particles (atoms, molecules, ions, etc.) as there are carbon atoms in 12 grams of the ${}_{6}C^{12}$ isotope.

Definition 2: A mole is defined as the amount of substance which contains Avogadro number (6.023×10^{23}) of particles.

Gram molecular weight. The molecular weight of a substance expressed in grams is known as gram molecular weight of the substance. The gram molecular weight of oxygen is 32g and that of sulphuric acid is 98g.

Molar volume. Volume occupied by one mole of any gas is called molar volume or gram molecular volume. It is 22.4L (or) 2.24×10^{-2} m^3at S.T.P. It contains 6.023×10^{23} molecules.

Equivalent mass of an element. Equivalent mass of an element is defined as the number of parts by mass of that element which can displace or combine with 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.46 parts by mass of chlorine or one equivalent mass of any other element. It is only a relative number and hence it does not have any units. When equivalent mass is expressed in gram, it is called gram equivalent mass.

Equivalent mass of an acid. Equivalent mass of an acid is the number of parts by mass of the acids which contains 1.008 parts by mass of replaceable hydrogen.

Basicity: Basicity of mineral acid is defined as the number of Replaceable hydrogen atoms present in one mole of the acid. Basicity of organic acid is defined as the number of carboxylic groups present in the acid.

Equivalent weight of base. Equivalent mass of the base is the number of parts by mass of the base required to neutralize one equivalent mass of an acid.

Acidity of a base. Acidity of hydroxide base is defined as the number of replaceable hydroxyl ions present in one mole of the base.

Equivalent mass of salt. Equivalent mass of a salt is the number of parts by mass of salt which reacts with one equivalent of mass of any other substance.

Equivalent weight of an oxidising agent. Equivalent weight of oxidizing agent is the number of parts by mass of it, which contains 8 parts by mass of available oxygen. Available oxygen means, oxygen capable of being utilised for oxidation.

Equivalent mass of a reducing agent. Equivalent weight of reducing agent is the number of parts by mass of it, which can be oxidized by 8 parts by mass of oxygen.

Normality of a solution. Normal solution is a solution, which contains one gram equivalent mass of the substance dissolved in one litre of the solution.

Law of volumetric analysis: When two solutions completely react with each other, the product of volume and normality of one solution will be equal to the product of volume and normality of the other solution.

Standard solution. In a titration, concentration of either the solution in the burette or in the conical flask should be exactly known. The solution whose concentration is exactly known is called the standard solution. A standard solution can be prepared by dissolving a known mass of the substance in a known volume of the solution.

SELF EVALUATION (T.B. PAGE 45 – 47)

I. CHOOSE THE CORRECT ANSWER.

1. The relationship between vapour density and molecular mass is ....
   (a) 2 × Density = Molecular mass (b) = Molecular mass
   (c) 2 × Vapour Density = Molecular mass (d) = Molecular mass
2. The volume occupied by 1 mole of the compound at STP is ....
   (a) mole (b) normality (c) vapour density (d) molar volume
3. How many moles are represented by 36 g of water?
   (a) 1 (b) 2 (c) 3 (d) 4
4. What is the mass of 4.48 X 10^{-2} m^3 of Methane gas at STP
   (a) 16g (b) 32g (c) 48g (d) 54g
5. A divalent element has 65.38 as its atomic mass. Its equivalent mass is ....
   (a) 32.69 (b) 65.38 (c) 130.76 (d) 16.35
6. The basicity of acetic acid CH_3COOH is ....
   (a) 2 (b) 1 (c) 3 (d) 0
7. A solution whose normality is known is called ....
   (a) Standard solution (b) Normal solution
   (c) Molar solution (d) None
8. For a monobasic acid, the molecular mass and equivalent mass are ....
   (a) same (b) different (c) reciprocal (d) their multiples
9. Gram molecular mass of nitrogen is ....
   (a) 28 (b) 14 (c) 7 (d) None
10. The number of molecules present in 17g of ammonia is ....
    (a) 6.023×10^{23} (b) 6.023 (c) 60.23×10^{23} (d) 6.023×10^{22}

Answers:

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

II. ANSWER THE FOLLOWING IN ONE OR TWO SENTENCES. (T.B. Page 46)

1. What do you mean by Avogadro Number?

It is defined as the number of atoms present in exactly 12 grams of $_6C^{12}$ isotope. It is denoted by N. It has a value of 6.023×10^{23}.

2. Define atomicity.

The number of atoms contained in one molecule of the element is called its atomicity.

3.Define equivalent mass of a reducing agent.

Equivalent weight of reducing agent is the number of parts by mass of it, which can be oxidized by 8 parts by mass of oxygen.

4.Define normality of a solution.

Normal solution is a solution which contains one gram equivalent mass of the substance dissolved in one litre of the solution.

5.State the law of volumetric analysis

When two solutions completely react with each other, the product of volume and normality of one solution will be equal to the product of volume and normality of the other solution.

$$V_1 \times N_1 = V_2 \times N_2$$

Where, V_1 and N_1 are the volume and normality of the first solution and V_2 and N_2 are the volume and normality of the second solution.

III. ANSWER IN DETAIL. (T.B. Page 46)

1. **State Avogadro's hypothesis. Apply it to deduce the relationship between vapour density and molecular mass.**

Avogadro's hypothesis. Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

Relation between Vapour Density & Relative Molecular Mass of a Gas:

Relative Molecular Mass: It is defined as the ratio of the mass of 1 molecule of the gas or vapour to the mass of 1 atom of hydrogen.

$$\text{Relative molecular mass of a gas} = \frac{\text{Mass of 1 molecule of the gas or vapour}}{\text{Mass of 1 atom of hydrogen}}$$

Vapour Density: It is defined as the ratio of the mass of a certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure.

$$\text{Vapour density (V. D.)} = \frac{\text{Mass of 1 volume of gas or vapour}}{\text{Mass of 1 volume of hydrogen}}$$

Applying Avogadro's Law,

$$\text{Vapour density (V.D)} = \frac{\text{Mass of 1 molecule of gas or vapour}}{\text{Mass of 1 molecule of hydrogen}}$$

Since hydrogen is diatomic,

$$\text{V.D.} = \frac{\text{Mass of 1 molecule of gas or vapour}}{\text{2 x Mass of 1 atom of hydrogen}}$$

Multiplying both sides by 2, we get

$$2 \times \text{V.D.} = \frac{\text{Mass of 1 molecule of gas or vapour}}{\text{Mass of 1 atom of hydrogen}}$$

$$\therefore 2 \times \text{V.D.} = \text{Relative molecular mass of the gas or vapour}$$

$$2 \times \text{Vapour Density} = \text{Molecular Weight.}$$

2.Explain Oxide formation method of determining equivalent mass of an element with an example.

Principle: This method is used to find out the equivalent mass of those metals (Magnesium, Zinc) which can easily form their oxides.

$$2Mg + O_2 \longrightarrow 2MgO$$

Procedure:

In general a known mass of a metal (m_1) whose equivalent mass is to be determined is heated in air or oxygen. The mass of metal oxide (m_2), formed is found.

In some cases (Copper, Lead, Tin etc.,) the metals could not be directly converted into their oxides. In such a situation, first the known mass of metal (m_1) is converted into its salt. Then the metallic salt formed is converted into its oxide by heating the salts. The mass of metal oxide formed is found (m_2).

From the mass of metal oxide and the mass of metal, the mass of oxygen that has combined with the metal is found.

The mass of oxygen that has combined with the metal $= m_2 - m_1$

$(m_2 - m_1)$ g of oxygen has combined with $= m_1$g of metal

$\therefore$ 8g of oxygen will combine with $= \left\{ \dfrac{m_1}{m_2 - m_1} \right\} \times 8$

By definition it refers to the equivalent mass of the metal.

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of metals}}{\text{Mass of oxygen that has combines with metal}} \times 8$$

IV. PROBLEMS FOR PRACTICE (T.B. PAGE 46).

1. **Calculate the molecular mass of a substance whose vapour density is 8.**

   Solution: The Vapour density of a substance (V.D) = 8

   Molecular weight = 2 × vapour density
   = 2 × V.D
   = 2 × 8
   = 16

   $\therefore$ The molecular mass of the substance is equal to 16.

2. **How many moles are contained in 32g of hydrogen?**

   Solution: Mass of Hydrogen = 32 g

   Molecular mass of hydrogen = 2 g

   $\therefore$ No of Moles = 16

   $\therefore$ 32 g of Hydrogen contains 16 moles of hydrogen.

3. **What is the mass of 10 moles of sodium hydroxide?**

   Solution: NaOH contains 10 no of moles.

   $\therefore$ Gram Molecular mass of NaOH = 40 g

   Mass of NaOH = No of moles × molecular mass
   = 10 × 40 = 400 g

   $\therefore$ 10 moles of NaOH contain 400 g of NaOH.

4. **What would be the volume of 14g of nitrogen at STP?**

   Solution: Mass of Nitrogen = 14 g

   Molecular mass of Nitrogen = 28 g

   No of Moles = 0.5 moles

   $\therefore$ 14 g of Nitrogen contains 0.5 moles.

   $\therefore$ Volume of the gas at S.T.P in Lit = No of moles of Nitrogen × 22.4
   = 0.5 × 22.4
   = 11.2 Litre.

5. **How many atoms are present in 10.8g of silver?**

   Solution:

   Gram atomic mass of silver = 108g

   1 mole of silver = 108 g

   108 g of silver = 1 mole of silver = 6.023×10^{23} atoms

   $\therefore$ 10.8 g of Silver contain = 0.1 mole.

   $$10.8 \text{ silver contain} = \frac{6.023 \times 10^{23}}{108} \times 10.8$$

   $= 6.023 \times 10^{22}$ atoms

6. **What is the mass of 1 atom of copper?**

   Solution:

   Atomic mass of Copper is = 63.546

$$\text{Mass of 1 atom of Copper} = \frac{\text{Atomic mass}}{6.023 \times 10^{23}} \text{ g}$$

$$= \frac{63.546}{6.023 \times 10^{23}}$$

$$\therefore \text{Mass 1 atom of Copper} = 10.46 \times 10^{-23} \text{ g}$$

7. In hydrogen displacement method, 0. 56 g of metal displaced 2.45×10^{-4} m^3 of hydrogen gas from dilute acid at 27°C and 1.00633×10^5 Nm^{-2}. The aqueous tension at 27^0C is 3.558×10^3 Nm^{-2}. Calculate the equivalent mass of the metal.

Solution:

The pressure of dry hydrogen displaced (P_1) = Pressure of moist hydrogen (P) - Aquoustension (P)

$= (1.00633 \times 10^5) – (3.558 \times 10^3)$

$= (100.633 \times 10^3) – (3.558 \times 10^3)$

$= (100.633 – 3.558) \times 10^3$

$= 97.075 \times 10^3$ Nm^{-2}

$\therefore$ Pressure of dry hydrogen P_1 = 9.7075×10^4 Nm^{-2}

The volume of hydrogen (V_1) displaced = 2.45×10^{-4} m^3

The room temperature (T_1) = 27 + 273

= 300 K

Standard temperature (P_0) = 1.013×10^5 Nm^{-2}

Standard temperature (T_0) = 273 K

$$\frac{P_1V_1}{T_1} = \frac{P_0V_0}{T_0}$$

$$V_0 = \frac{P_1V_1}{T_1} \times \frac{T_0}{P_0}$$

$$= \frac{9.7075 \times 10^4 \times 2.45 \times 10^{-4}}{300 \text{ K}} \times \frac{273}{1.013 \times 10^5}$$

$$= 21.365 \times 10^{-5} \text{ m}^3$$

At STP 21.365×10^{-5} m^3 of hydrogen displaced by 0.56 g of metal.

$\therefore$ At STP 1.12×10^{-2} m^3 of hydrogen will be displaced by

$$= \frac{0.56}{21.365 \times 10^{-5}} \times 1.12 \times 10^{-2}$$

$$= 0.029 \times 10^3 = 29$$

$\therefore$ Equivalent mass of the metal = 29.36

8. A metal oxide has 60 % of metal. Calculate the equivalent mass of the metal. ***(Hint: Assume the weight of metaloxide as 100%).***

Solution: Mass of Metal = 60% = 0.6 g

Mass of metal oxide = 100% = 1 g

Mass of oxygen = (1 –0.6) = 0.4 g

$$\text{Equivalent mass of metal} = \frac{\text{Mass of metal}}{\text{Mass of oxygen that has combine with metal}} \times 8$$

$$= \frac{0.6}{} \times 8$$

0.4

Equivalent mass of metal = 12.

9.A metal chloride has 60.66% of chlorine. Calculate the equivalent mass of the metal. ***(Hint: Assume the weight of metalchloride as 100%).***

Solution:

Mass of Metal chloride = 100 grams

Mass of chlorine = 60 grams

Mass of metal = 100 – 60.66 = 39.34 g

$$\text{Equivalent mass of the metal} = \frac{39.34 \times 35.46}{60.66} = 23$$

Equivalent mass of the metal = 23.

10. 1.26g of oxalic acid crystals are dissolved in 500 ml of the solution. Calculate its normality.

Solution:

Equivalent mass of oxalic acid = 63

$$\text{Normality} = \frac{\text{Mass in grams per litre of solution}}{\text{Equivalent mass}}$$

1.26 grams of oxalic acid dissolved in 500 ml = 2.53 grams of oxalic acid dissolved in 1000 ml

$$= \frac{2.52}{63}$$

= 0.04 N

∴ The Normality of oxalic acid = 0.04 N.

11.60g of crystalline sodium hydroxide is dissolved in 2 dm^3 of a solution. Calculate its normality.

Solution:

Equivalent mass of NaOH = 40

60 g of NaOH dissolved in 2 lit = 30 g of NaOH dissolved in 1 lit

$$\text{Normality} = \frac{\text{Mass in grams per litre of solution}}{\text{Equivalent mass}}$$

$$= \frac{30}{40}$$

= 0.75 N

∴ Normality of NaOH = 0.75 N.

OTHER IMPORTANT QUESTIONS & ANSWERS

I. CHOOSE THE CORRECT ANSWER:

1. Which of the following equation represent the correct relationship?
   (a) equivalent weight = atomic weight x valency (b) valency = equivalent weight x atomic weight
   (c) atomic weight = equivalent weight x valency (d) $\text{atomic weight} = \frac{\text{equivalent weight x valency}}{2}$
2. If an element has an atomic weight of 24 and valency of 2, its equivalent weight is,
   (a) 24 (b) 48 (c) 12 (d) 57.
3. The number of silicon atoms present in 0.778 mol of silicon
   (a) 4.69×10^{23} (b) 7.74×10^{23} (c) 6.02×10^{23} (d) 8.77×10^{24}
4. The molecular weight of $(NH_4)_2SO_4$ is
   (a) 114 (b) 146 (c) 132 (d) 70
5. Which of the following contains largest number of atoms?
   (a) 2.0 mol of H (b) 9 g of CH_4 (c) 18.0 g H_2 (d) 10.0 g Cl_2
6. Which of the following contains 6.02×10^{33} molecules
   (a) 12 g CH_4 (b) 6 g H_2O (c) 5 g C_2H_6 (d) 17 g NH_3

7. The volume in dm^3 occupied by 1 mole of a gas at STP is
(a) 12.0 (b) 22.4 (c) 24.0 (d) 100

8. Which of the following relationship is correct?
(a) vapour density = 2 x relationship weight (b) vapour density = 2 x molecular weight
(c) 2 x vapour density = molecular weight (d) molecular weight x vapour density = 2

9. The number of molecules of HC*l* present in 28.5 g of the sample is:
(a) 0.425×10^{23} (b) 4.71×10^{23} (c) 2.125×10^{23} (d) 0.2125×10^{23}

10. How many mole of helium gas occupy 22.4 litres at S.T.P?
(a) 0.11 (b) 2.090 (c) .1.0 (d) 1.11

11. One mole of CO_2 contains
(a) 6.02×10^{23} atoms of C (b) 6.02×10^{23} atoms of O
(c) 18.1×10^{23} (d) 3 gram $\times 10^{23}$ molecules of CO_2

12. 2 moles of H atoms at STP occupies a volume of:
(a) 11.2 litre (b) 44.8 litre (c) 2 litre (d) 22.4 litre.

13. Vapour density of a volatile substance is 4.It's molecular weight would be:
(a) 8 (b) 2 (c) 64 (d) 128

14. Equivalent weight of an element is equal to
(a) atomic weight x valency (b) atomic weight ÷ valency
(c) valency ÷ atomic weight (d) molecular weight x valency

15. One gram of hydrogen at S.T.P occupies a volume equal to
(a) 22.4 litres (b) 10 litres (c) 11.2 litres (d) 20 litres.

16. If the molecular weight of NaOH is 40, its equivalent weight is
(a) 40 (b) 98 (c) 198 (d) 147

17. The number of atoms in 24 g of magnesium is close to
(a) 24 (b) 2×10^{20} (c) 10^{20} (d) 6.02×10^{23}

18. The total number of atoms present in 18 ml of water (density of water is 1 g ml^{-1}) is
(a) 6.02×10^{23} (b) 6.02×10^{22} (c) 6.02×10^{24} (d) 6.02×10^{25}

19. 18 g of water contain
(a) 1 g atom of hydrogen (b) 2 g atoms of hydrogen
(c) 3 g atoms of hydrogen (d) None of the above

20. The mass of an atom of oxygen is
(a) 16 amu (b) $16 / 6.023 \times 10^{23}$ g (c) $32 / 6.023 \times 10^{23}$ g (d) $1 / 6.023 \times 10^{23}$ g

21. 1 mole of methane (CH_4) contains
(a) 6.02×10^{23} atoms of H (b) 14 gram atoms of hydrogen
(c) 1.81×10^{23} molecules of methane (d) 3.0 g of carbon

22. Volume occupied by 64g of SO_2 in S.T.P .is …………
(a) 22.4 L (b) 11.2 L (c) 44.8 L (d) 32 L

23. The atomicity of hydrogen is
(a) 1 (b) 3 (c) 2 (d) 4

24. At S.T.P 1 volume of gas combined 1 volume of another gas to forms…………
(a) 2 volumes of gas (b) 1 volume of gas (c) 3 volumes of gas (d) volume of gas

25. The ratio of mass of certain volume of gas to the mass of same volume of hydrogen at S.T.P is called
(a) Molecular weight (b) vapour density (c) mole (d) atomic mass

26.The molecular mass of oxygen is 32, the vapour density is …………
(a) 32 (b) 16 (c) 64 (d) 96

27. 1 mole of nitrate ions contains………………. ions of NO^-_3
(a) 10 (b) 6.023 (c) 6.023×10^{22} (d) 6.023×10^{23}

28. The gram atomic mass of the substance is 3.015, the mass of 1 atom ……………
(a) 0.5×10^{-23} (b) 0.2×10^{-23} (c) 0.6×10^{-23} (d) 0.1×10^{-23}

29. 1.204×10^{23} atoms of sodium contains
(a) 0.2 moles (b) 0.3 moles (c) 0.4 moles (d) 0.1 moles

30. Equivalent mass of magnesium is
(a) 14.06 (b) 13.06 (c) 12.06 (d) 15.06

31. The number of replaceable hydrogen atoms present in 1 mole of the acid is called……..
(a) Basicity (b) Acidity (c) normality (d) equivalent mass

32. The molecular mass of sulphuric acid is 98, the equivalent mass is ………….
(a) 98 (b) 40 (c) 186 (d) 49

33. The volume required when 20ml of 12 N HC*l* react with 6 N NaOH is ……….
(a) 20 ml (b) 30 ml (c) 15 ml (d) 40 ml

34.The strength required when 15 ml of 1 N sodium hydroxide neutralized by 10 ml of HCl
(a) 10 N (b) 0.5 N (c) 1.5 N (d) 1 N

35. How many number of molecules contain 2 moles of oxygen molecule
(a) 6.023×10^{23} (b) 6.023×10^{24} (c) 12.046×10^{23} (d) 18.023×10^{23}

36. 0.5 moles of gas contains …………… volume at S.T.P
(a) 22.4 L (b) 44.8 L (c) 1.2 L (d) 11.2 L

37. The equivalent mass of acetic acid is 60. The molecular mass of acetic acid is ………….
(a) 30 (b) 60 (c) 120 (d) 15

38. The gram atomic weight of carbon is …………..
(a) 6g (b) 12g (c) 16g (d) 18g
39. The atomic weight of an element expressed in gram is known as………
(a) gram atomic weight (b) gram molecular weight
(c) gram equivalent mass (d) None of these
40. The ratio between mass of one atom of the element to the mass of one atom of hydrogen is called…….
(a) relative atomic mass (b) relative molar mass
(c) relative equivalent mass (d) electrochemical equivalent
41. Vapour density is equal to………
(a) 2 x molecular weight (b) molecular weight
(c) $\frac{\text{Atomic weight}}{2}$ (d) $\frac{\text{Molecular weight}}{2}$
42. The ratio of the mass of a certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure is called………..
(a) density (b) vapour density (c) molecular formula (d) None of these
43. The ratio of the mass of one molecule of the gas or vapour to the mass of one atom of hydrogen is ………..
(a) molecular mass (b) relative atomic mass
(c) relative moelcular mass (d) vapour density.
44. A mole is defined as the amount of substance which contains……. of particles.
(a) Avogadro number (b) 6.023×10^{23} (c) 6.023×10^{10} (d) (a) and (b)
45. A mole is the amount of substance which contains as many particles as there are carbon atoms in………
(a) 12 grams of the C^{12} isotope (b) 12 grams of the C^{13} isotope
(c) 12 grams of the C^{14} isotope (d) all of these
46. One mole of oxygen atoms contain………
(a) 6.023×10^{23} atoms of oxygen (b) 6.023×10^{23} moleculses of oxygen
(c) 0.6023×10^{23} atoms of oxygen (d) $2 \times 6.023 \times 10^{23}$ atoms of oxygen
47. 6.023×10^{23} molecules of oxygen is equal to………..
(a) 0.5 mole of oxygen molecule (b) 0.5 mole of oxygen atom
(b) 1.0 mole of oxygen atom (c) 1.0 mole of oxygen molecule
48. One mole of nitrate ions = ………….
(a) 6.023×10^{23} ions of NO_3^- (b) 6.023×10^{23} atoms of NO_3
© 6.023×10^{23} molecules of nitrite (d) all of these
49. One mole always contain 6.023×10^{23} particles. Thr term particles refer to ………
(a) atoms (b) molecules (c) ions (d) all of these
50. The number of atoms contained in one molecule of the element is called its……..
(a) atomicity (b) atomic mass (c) atomic number (d) molar volume
51. Two molecules of hydrogn chloride contain……..of hydrogen and…….. of chlorine
(a) 1 and 2 (b) 2 and 2 (c) 2 and 1 (d) 1 and 1
52. Hydrogen molecule is……..
(a) mono atomic (b) triatomic (c) diatomic (d) poly atomic
53. What is the atomicity of hydrogen?
(a) 1 (b) 2 (c)3 (d) None
54. Atomicity of chlorine is ………
(a) 2 (b) 1 (c) 3 (d) 0
55. Nitrogen has atomicity………….
(a) 1 (b) 3 (c) 2 (d) 0
56. Which of the following gases have atomicity two?
(a) H_2,N_2, He (b) N_2,$C1_2$,Ar (c) Na, H_2, N_2 (d) H_2, Cl_2, O_2, N_2
57. Atomicity of oxygen is
(a) 1 (b) 0 (c) 2 (d) 3
58. Which has atomicity equal to three?
(a) Oxygen (b) Ozone (c) Nitrogen (d) Hydrogen
59. The gram atomic weight of oxygen is………
(a) 32g (b) 12g (c) 16g (d) 18g
60. The relative molar mass is equal to……..
(a) $\frac{\text{mass of one atom of the element}}{\text{mass of one atom of hydrogen}}$ (b) $\frac{\text{mass of one molecule of the element}}{\text{mass of one atom of hydrogen}}$
(c) $\frac{\text{mass of one molecule of the element}}{\text{mass of one molecule of hydrogen}}$ (d) None of these
61. When the molecular weight of a substance expressed in grams is known as………
(a) gram molecular weight (b) gram atomic weight
(c) gram equivalent mass (d) all of these

62. The gram molar mass of oxygen is…….
(a) 8g (b)16g (c) 24g (d) 32g
63. The gram molecular weight of sulphuric acid is……….
(a) 49g (b) 98g (c) 32g (d) 64g
64. How will you express the quantity of substance in general?
(a) Gram (b) Kilogram (c) Mole (d) Mg
65. When mole concept is used, it is necessary to specify the………
(a) kind of particles like atoms, molecule or ions (b) valency
(c) atomic weight (d)atomic number
66. If the kind of particles is not expressed in the mole concept generally it is assumed to be…………
(a) atoms (b) ions (c) molecules (d) all tof these
67. One mole of chlorine atom represents…………
(a) 6.023×10^{23}molecules (b) 6.023×10^{23}atoms
(b) 6.023×10^{23} ions (d) all of these
68. One mole of chlorine means……….
(a) 6.023×10^{23}molecules of Cl_2 (b) 6.023×10^{23} atoms of Cl
(c) 6.023×10^{23} chloride ions (d) all of these
69. The number of atoms present exactly in12 grams of $_6C^{12}$ isotope is defined as ……
(a) Avogadro's hypothesis (b) Berzelius hypothesis
(c) Avogadro's number (d) Oxidation number
70. Avogadro number is denoted by the symbol ……….
(a) A (b) Z (c) M (d) N
71. The product of number of moles and 6.023×10^{23} gives………….
(a) number of atoms (b) number of molecules
(c) both (a) and (b) (d) None of these
72. How many moles are represented by 92 g of sodium?
(a) 1 (b) 2 (c) 3 (d) 4
73. Three moles of oxygen atoms equal to ……….
(a) 16g (b) 32g (c) 48g (d) None of these
74. How many moles are there in 3 grams of hydrogen?
(a) 1.5 moles (b) 2 moles (c) 2.5 moles (d) 0.5 mole
75. Calculate the number of moles present in 22g of CO_2
(a) 0.25 (b) 0.50 (c) 1.0 (d) 2.5
76. 0.5 mole of NaOH = ……….. g.
(a) 40 (b) 10 (c) 20 (d) 30
77. Number of moles can be calculated using the formula………..
(a) $\dfrac{\text{atomic mass}}{\text{mass}}$ (b) $\dfrac{\text{mass}}{\text{atomic mass}}$ (c) $\dfrac{\text{mass}}{\text{molecular mass}}$ (d) both (b) and (c)
78. Two moles of substance contains……… number of molecules.
(a) 6.023×10^{23} (b) $2 \times 6.023 \times 10^{23}$ (c) $6.023 \times 10^{23} \times 3$ (d) $\dfrac{6.023 \times 10^{23}}{2}$
79. 3.0115×10^{23} molecules of water are present in …………. of water.
(a) 18g (b) 36 g (c) 4.5 (d) 9g
80.22 g of CO_2 will contain………. Numbe of molecules.
(a) 6.023×10^{23} (b) 3.0115×10^{23} (c) 1.505×10^{23} (d) None of these.
81. Calculate the number of molecules present in 48g of ozone.
(a) 6.023×10^{23} molecules (b) Avogadro number of molecules
(c) both (a) and (b) (d) None of these
82. The gram molecular or molar mass of ozone (O_3) is………..
(a) 32g (b)16g (c) 48g (d)44g
83. The gram molecular mass of water is……..
(a) 16g (b) 18g (c) 32g (d) 48g
84. What is the mass of 0.2 mole of oxygen molecule?
(a) 12.8g (b) 32g (c) 16g (d) 6.4g
85. Calculate the mass of 1.5 mole of CO_2.
(a) 44g (b) 66g (c) 110g (d) 88g
86. Calculate the mass of 0.5 mole of copper. (at.wt.of Cu = 63.54) =
(a) 63.54g (b) 15.89g (c) 31.77g (d) 127.08g
87. 1.10 moles of nitrogen atoms = ………….
(a) 30.8g (b) 14g (c) 28g (d) 15.4 g
88. In general, the mass of one mole of the substance is equal to its……..
(a) gram molecular mass (b) equivalent mass
(c) electro chemical equivalent (d) all of these
89. What is the gram molecular mass of CO_2?
(a) 440 g (b) 44g (c) 12g (d) 32g

90. $\dfrac{\text{Gram molecular mass}}{\text{Avogadro number}}$ = …………

(a) mass of one atom (b) mass of one molecule

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

91. The mass of one atom of any element is calculated using the formula…………..

(a) $\dfrac{\text{Atomic number}}{\text{Avogadro number}}$ (b) $\dfrac{\text{Avogadro number}}{\text{Gram atomic mass}}$

(c) Gram atomic mass x Avogadro number (d) $\dfrac{\text{Gram atomic mass}}{\text{Avogadro number}}$

92. One gram atomic mass of any element contains………… Of atoms in general.

(a) Avogadro number (b) definite

(c) indefinite (d) none

93. Avogadro number has a value of ………..

(a) 6.023×10^{23} (b) 60.023×10^{23} (c) 0.6023×10^{23} (d) 6.023×10^{25}

94. What is the molecular of glucose? (Molecular formula $C_6H_{12}O_6$)

(a) 90 (b) 180 (c) 360 (d) 250

95. 45 g of glucose contains …………… number of molecules.

(a) 6.023×10^{23} (b) 3.011×10^{23} (c) 1.505×10^{23} (d) None of these

96. Two gram molecular mass of ammonia (NH_3) is …………

(a) 17 g (b) 34 g (c) 8.5 g (d) 51 g

97. The gram molar or molecular mass of sucrose is ………..

(a) 180 g (b) 342 g (c) 260 g (d) 90 g

98. The volume of 22.4 L of any gas at STP is called …………..

(a) vapour density (b) molar volume (c) normality (d) molality

99. The volume of one mole of any gas at STP occupies ………….

(a) 2.24×10^{-2} m^3 (b) 2.24 L (c) 22.4×10^{-2} m^3 (d) None of these

100. One mole of any gas at STP equal to / occupies ………..

(a) 2.24 L (b) 2.24×10^{-2} m^3 (c) both (d) None of these

101. Which has been displaced from their compounds in displacement methods?

(a) Hydrogen (b) Oxygen (c) Metal (d) All of these

102. The equivalent mass of an element is determined from the ………….. displaced.

(a) amount of hydrogen (b) amount of oxygen (c) amount of metal (d) all of these

103. Which metals readily displace hydrogen gas from an acid?

(a) Na (b) Zn and Al (c) Mg (d) All of these

104. Equivalent mass of the metal = ………………

(a) $\dfrac{\text{Mass of the metal}}{V_0 \text{ m}^3}$ (b) $\dfrac{\text{Mass of the metal} \times 1.12 \times 10^{-2} \text{ m}_3}{V_o \text{ m}^3}$

(c) $\dfrac{V_o \text{ m}^3}{\text{Mass of the metal}}$ (d) None of these

105. Laboratory temperature (T_1), pressure (P_1) and volume (V_1) is converted into STP values by ………

(a) $\dfrac{P_1 V_1}{T_1} = \dfrac{P_0 V_0}{T_0}$ (b) $P_1 V_1 = P_0 V_0$ (c) $\dfrac{1}{T_1} = \dfrac{1}{T_0}$ (d) $\dfrac{V_1}{T_1} = \dfrac{V_0}{T_0}$

106. The pressure of dry hydrogen displaced (P_1) is …………

(a) P – p (b) p – P (c) P x p (d) P / p

107. Combination method involves ………….

(a) oxide method (b) chloride method (c) both (a) and (b) (d) none of these

108. 22.4 L or 2.24 x 10-2 m3 of any gas at STP will contain ……………….. molecules.

(a) 6.023 x 10^{23} (b) Avogadro number (c) both (a) and (b) (d) None of these

109. The volume of the gas at STP in litre = ………………..

(a) $\frac{\text{No. of moles of a gas}}{\text{22.4 litre}}$ (b) $\frac{\text{22.4 litre}}{\text{No. of moles of a gas}}$

(c) No. of moles of a gas x 22.4 litre (d) all of these

110. What is the volume of 0.25 moles of CO_2?

(a) 22.4 L (b) 11.2 L (c) 5.6 L (d) 2.24 x 10^{-2} m^3

111. What is the volume of 0.5 mole of CO2?

(a) 11.2 L (b) 1.12 x 10^{-2} m^3 (c) both (a) and (b) (d) None of these

112. Elements having varying valency have …………… equivalent weight.

(a) same (b) different (c) none (d) constant

113. Equivalent mass is only a ……………..

(a) relative number (b) number (c) mass (d) none

114. The equivalent mass has ………………..

(a) units (b) dimensions (c) both (a) and (b) (d) no units

115. When equivalent mass is expressed in gram, it is called …………..

(a) atomic mass (b) gram equivalent mass (c) gram molar mass (d) molar mass

116. Equivalent mass of an element is defined as the number of parts by mass of that element which can displace or combine with ……………….

(a) 1.008 parts by mass of hydrogen (b) 8 parts by mass of oxygen

(c) 35.46 parts by mass of chlorine (d) all of these

117. The equivalent weight of an element = …………………..

(a) Atomic weight x Valency (b) $\frac{\text{Valency}}{\text{Equivalent weight}}$

(c) both (a) and (b) (d) $\frac{\text{Atomic weight}}{\text{Valency}}$

118. Atomic weight itself equal to equivalent weight in the case of ……………. Elements.

(a) univalent (b) divalent (c) trivalent (d) None of these

119. Equivalent weight of copper in cuprous oxide (valency of Cu is one) is ………….

(a) 31.75 (b) 63.5 (c) both (a) and (b) (d) None

120. Equivalent weight of copper in cupric oxide (CuO) (valency of Cu is two) is ………………

(a) 31.75 (b) 63.5 (c) 108 (d) 127

121. The equivalent mass of an element could be determined by ……………….. methods.

(a) displacement (b) combination (c) separation (d) both (a) and (b)

122. Normality = ……………….

(a) $\frac{\text{Mass in gram per litre of solution}}{\text{Equivalent mass}}$ (b) $\frac{\text{Mass}}{\text{Valency}}$

(c) $\frac{\text{Molar mass}}{\text{Valency}}$ (d) Equivalent mass x Valency

123. A standard solution containing 1 / 10^{th} of the gram equivalent of the substance is called …………….

(a) decinormal solution (b) 1 / 10N (c) 0.1 N (d) all of these

124. Two normal (2N) solutions …………….. of the substance in a standard solution.

(a) one gram equivalent (b) two gram equivalent (c) one mole (d) two moles

125. A standard solution containing 1 gram equivalent of the substance is called …………….. solution.

(a) 1 N or one normal (b) 2N (c) N / 10 (d) 0.1 N

126. Volumetric analysis depends on ………………

(a) temperature (b) catalyst

(c) volumes of solutions of interacting substances (d) all of these

127. A trivalent element has 27 as its atomic mass. Its equivalent mass is ………..

(a) 27 (b) 13.5 (c) 9 (d) None

128. The mathematical form of volumetric law is ……………….

(a) $V_1 N_1 = V_2 N_2$ (b) $\frac{V_1}{N_1} = \frac{V_2}{N_2}$ (c) $V_1 = V_2$ (d) $N_1 = N_2$

129. The amount (mass) of the substance present in one litre of the solution is found out by …………..

(a) $\frac{\text{Normality}}{\text{Equivalent mass}}$ (b) Normality x Equivalent mass

(c) $\frac{\text{Equivalent mass}}{\text{Normality}}$ (d) $\sqrt{N \times \text{Equivalent mass}}$

130. The volumes of solutions that are reacting with each other are determining by the process called …….

(a) hydration (b) hydrolysis (c) titration (d) electrolysis

131.Equivalent mass of the salt is …………….

(a) $\frac{\text{Molar mass of the salt}}{\text{Atomicity}}$ (b) Valency x 2

(c) $\frac{\text{Molar mass of the salt}}{\text{Valency of the metal in the salt}}$ (d) $\frac{\text{Atomicity}}{2}$

132. The molecular mass of $CuCl_2$ is ……………. (Cu = 63.5)

(a) 134.42 (b) 63.5 (c) 35.46 (d) 71.0

133. Equivalent mass of $CuCl_2$ is ………………..

(a) 134.42 (b) 67.21 (c) 40 (d) 71

134. The total valency of sodium in Na_2SO_4 ……………….

(a) 1 (b) 0 (c) 2 (d) None

135. Volumetric analysis involves …………..

(a) estimation of a substance (b) characterization of a substance

(c) identification of an acid radical (d) identification of a basic radical

136. We can estimate the substance by ………….. volumetric analysis.

(a) neutralisation (b) oxidation or reduction (c) precipitation (d) all of these

137. 0.09 g of an element displaces 56 x 10^{-6} m^3 of oxygen at STP. The equivalent mass of the element ..
(a) 9. (b) 18. (c) 27. (d) 36.

138. One mole of chlorine when reacted with a metal gave 111 g of its chloride. The equivalent mass of the metal is …………
(a) 12. (b) 20. (c) 55.5. (d) 111.

139. The mass in gram of 22.4 dm^3 of any gas at NTP is ………..
(a) equivalent mass. (b) atomic mass
(c) vapour density. (d) gram molecular mass.

140. Which of the following will contain the same number of atoms as 20 g of calcium?
(a) 24 g of Mg. (b) 12 g or C (c) 24 g of C (d) 12 g of Mg.

141. Avagadro number of helium atoms weigh……….
(a) 1.00 g. (b) 4.00 g. (c) 8.00 g. (d) $4 \times 6.02 \times 10^{23}$

142. The largest number of molecules are in………..
(a) 28 gCO_2. (b) 36 g H_2O. (c) $46gC_2H_5OH$. (d) 54g N_2O_5.

143. 4.25 g of ammonia is equal to ………..
(a) 0.25 mole (b) 1 mole. (c) 1.5 mole. (d) 0.5 mole.

144. The number of molecules in 32g of oxygen is …………..
(a) 3.2×10^{16}. (b) 6.02×10^{23}. (c) 3.2×10^{23} (d) 6.02×10^{18}.

145. The molecular mass of a compound KCl is 74.5. The equivalent mass of the metal will be……
(a) 74.5. (b) 110. (c) 37.2. (d) 39.

146. How many molecules are present in one gram of hydrogen?
(a) 6.023×10^{22} (b) 6.023×10^{23} (c) 3.015×10^{23}. (d) 3.015×10^{-12}.

147. The weight of one calcium atom (at mass = 40) is ……..
(a) 40g (b) 6.02×10^{-23}g. (c) 6.64×10^{-23}g. (d) 6.02×10^{23}g.

148. The number of moles in 1 kg of Fe_2O_3 is ………
(a) 1. (b) 2. (c) 5.5. (d) 6.26.

149. The volume at NTP of hydrogen produced by 12 g Mg (at. Wt. 24)………
(a) 2.24×10^{-2} m^3. (b) $1.12 \times 10^{-2} m^3$. (c) 44.8 dm^3 (d) 6.1 dm^3.

150. The mass of 2.24 $dm^3$of a gas under standard conditions is 2.8 g. Its molar mass is…..
(a) 28. (b) 14. (c) 42. (d) 56.

151. The vapour density of a gas is 11.2. The volume occupied by 11.2 $dm^3$of the gas at NTP is …….
(a) 1 dm^3. (b) 11.2 dm^3. (c) 22.4 dm^3. (d)10dm^3.

152. The oxidation number of chlorine in ClO_4^- is ……….
(a) +7. (b) –7. (c) +4. (d) –4.

153. The oxidation number of nitrogen is highest in………..
(a) No_2. (b) HNO_2. (c) N_2O_5. (d) NO.

154. Which one of the following is incorrect?
(a) During oxidation, the oxidation number of an element increases.
(b) During reduction the oxidation number of an element decreases.
(c) In it reactions when MnO_4^- is converted to Mn^{+2}, the oxidation number of manganese decreases.
(d) The oxidation number of iron in ferrous sulphate is +3.

155. The oxidation number of copper is……..
(a) 0. (b) +1. (c) –1. (d) +2.

156. The number of equivalents ferrous sulphate that is oxidised by 3.16 g of acidified $KMnO_4$ is ……….
(a) one. (b) 0.1. (c) 0.2 (d) 5.

157. Which of the following containing the same number of moles?……
(a) 49 g of H_2SO_4. (b) 100 g. of $CaBr_2$. (c) 75 g of NaI. (d) all.

158.0.35 g of a volatile liquid is displaced $8.65 \times 10^{-5} m^3$ of air at NTP in a Victor Meyer's apparatus. The molecular mass of the liquid is……..
(a) 90.64. (b) 45.32 (c) 118.20 (d) 80.00.

159. The molecular weight of NO_2 is 46. Its density in gdm^{-3}will be…….
(a) $\frac{46}{22.4}$. (b) $\frac{46}{22400}$. (c)46×22.4. (d)$46 \times \frac{22400}{760}$

160. 40ml of 0.1 N HCl will react with 20 ml of ………
(a) 0.1 N NaOH. (b) 0.05 N NaOH.
(c) 0.2 N NaOH. (d) 0.15 N NaOH.

161. Phosphoric acid is a tribasic acid. It has a molecular weight of 98. Its equivalent weight is……..
(a) 294. (b) 98. (c) 32.6. (d) 196.

162. The equivalent weight of a trivalent element is 9. Its atomic weight is …….
(a) 18. (b) 81. (c) 9. (d) 27.

163.The equivalent weight of acetic acid CH_3COOH is ……….
(a) 60. (b) 30. (c) 15. (d) 120.

164. The atomic weight of iron is 55.8. Its equivalent weight in ferrous chloride is……
(a) 27.9. (b) 55.8. (c) 63.4. (d) 18.6.

165. Equivalent weight of an oxidizing agent or a reducing agent is equal to the ratio of its formula weight to
(a) valency. (b) atomic weight.
(c) change in oxidation number. (d) acidity.

Answers:

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

14. (b) 15. (a) 16. (a) 17. (d) 18. (a) 19. (b) 20. (a) 21. (a) 22. (a) 23. (c) 24. (a) 25. (b) 26. (b)

27. (d) 28. (a) 29. (a) 30. (c) 31. (a) 32. (d) 33. (d) 34. (c) 35. (c) 36. (d) 37. (b) 38. (b) 39. (a)

40. (a) 41. (d) 42. (b) 43. (c) 44. (d) 45. (a) 46. (a) 47. (d) 48. (a) 49. (d) 50. (a) 51. (b) 52. (c) 53. (b)

54. (a) 55. (c) 56. (d) 57. (c) 58. (b) 59. (c) 60. (b) 61. (a) 62. (d) 63. (b) 64. (c) 65. (a) 66. (c) 67. (b) 68. (a)

69. (c) 70. (d) 71. (c) 72. (d) 73. (c) 74. (a) 75. (b) 76. (c) 77. (d) 78.(b) 79. (d) 80. (b) 81. (c) 82. (c) 83. (b)

84. (d) 85. (b) 86. (c) 87. (d) 88. (a) 89. (b) 90. (b) 91. (d) 92. (a) 93. (a) 94. (b) 95. (c) 96. (b) 97. (b) 98. (b) 99. (a) 100. (b) 101. (d) 102. (d) 103. (d) 104. (b) 105. (a) 106. (a) 107. (c) 108. (c) 109. (c) 110. (c) 111. (c) 112. (b) 113. (a) 114. (d) 115. (b) 116. (d) 117. (d) 118. (a) 119. (b) 120. (a) 121. (d) 122. (a) 123. (d) 124. (b) 125. (a) 126. (c) 127. (c) 128. (a) 129. (b) 130. (c) 131. (c) 132. (a) 133. (b) 134. (c) 135. (a) 136. (d) 137. (a) 138. (b) 139. (d) 140. (b) 141. (d) 142. (b) 143. (a) 144. (b) 145. (a) 146. (c) 147. (c) 148. (d) 149. (b) 150. (a) 151. (b) 152. (a) 153. (c) 154. (d) 155. (a) 156. (b) 157. (d) 158. (a) 159. (a) 160. (c) 161. (c) 162. (d) 163. (a) 164. (a) 165. (c).

II. ANSWER THE FOLLOWING IN ONE OR TWO SENTENCES:

1.State Avogadro's hypothesis.

Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

2.Define vapour density.

Vapour Density is defined as the ratio of the mass of a certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure.

3. Define molecular mass.

The relative molecular mass of an element or a compound is the mass of one molecule of the element or compound compared with the mass of one atom of hydrogen taken as one unit.

4. Define atomic weight.

The relative atomic mass of an element is the mass of one atom of the element compared with the mass of one atom of hydrogen taken as one unit.

5. Define gram atomic weight of an atom.

The atomic weight of an element expressed in grams is known as the gram atomic weight (or gram atom) of the element.

For example, Gram atomic weight of nitrogen = 14 g

Gram atomic weight of oxygen = 16 g.

6. Define mole.

Definition 1: The mole is the amount of substance which contains as many particles (atoms, molecules, ions, etc.) as there are carbon atoms in 12 grams of the ${}_6C^{12}$ isotope.

Definition 2: A mole is defined as the amount of substance which contains Avogadro number (6.023×10^{23}) of particles.

7. Define gram molecular weight.

The molecular weight of a substance expressed in grams is known as gram molecular weight of the substance.

The gram molecular weight of oxygen is 32g and that of sulphuric acid is 98g.

8. Define molar volume.

Volume occupied by one mole of any gas is called molar volume or gram molecular volume. It is 22.4L (or) 2.24×10^{-2} m^3at S.T.P. It contains 6.023×10^{23} molecules.

9. Define equivalent mass of an element.

Equivalent mass of an element is defined as the number of parts by mass of that element which can displace or combine with 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.46 parts by mass of chlorine or one equivalent mass of any other element. It is only a relative number and hence it does not have any units

10. Give the relationship between equivalent weight and atomic weight?

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

11. Atoms of the same element have different equivalent weights – Why?

Elements having varying valency have different equivalent weight.

For example, Equivalent weight of copper in cuprous oxide

Cu_2O is 63.5. (here, valency of Cu is one).

Equivalent weight of copper in cupricoxide,

CuO is = 31.75 (here, valency of Cu is 2).

12. Define equivalent mass of an acid.

Equivalent mass of an acid is the number of parts by mass of the acids which contains 1.008 parts by mass of replaceable hydrogen.

$$\text{Equivalent weight of the Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the Acid}}$$

13. Define the term Basicity.

Basicity of mineral acid is defined as the number of Replaceable hydrogen atoms present in one mole of the acid. Basicity of organic acid is defined as the number of carboxylic groups present in the acid.

14. Define equivalent weight of base.

Equivalent mass of the base is the number of parts by mass of the base required to neutralize one equivalent mass of an acid.

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

15. Define acidity of a base.

Acidity of hydroxide base is defined as the number of replaceable hydroxyl ions present in one mole of the base.

16. Define equivalent mass of salt.

Equivalent mass of a salt is the number of parts by mass of salt which reacts with one equivalent of mass of any other substance.

$$\text{Equivalent weight of the Salt} = \frac{\text{Molecular mass of the Salt}}{\text{Valency of the metal in the salt}}$$

17. Define equivalent weight of an oxidising agent.

Equivalent weight of oxidizing agent is the number of parts by mass of it, which contains 8 parts by mass of available oxygen. Available oxygen means, oxygen capable of being utilised for oxidation.

18. What is called a standard solution?

In a titration, concentration of either the solution in the burette or in the conical flask should be exactly known. **The solution whose concentration is exactly known is called the standard solution.**

A standard solution can be prepared by dissolving a known mass of the substance in a known volume of the solution.

19. What is gram equivalent weight?

The equivalent weight of a substance is mere a number. It has no unit. If it is expressed in grams, it is known as gram equivalent weight.

20. Give the importance of equivalent weight?

(i) If the equivalent weight of an element and its valency are known, the atomic weight of an element can be calculated.

(ii) The principle of volumetric analysis is based on the law of equivalents which states that one equivalent weight of one substance reacts with one equivalent weight of another substance.

21. How do you distinguish between standard solution and a normal solution?

A standard solution is one whose normality is known. A normal solution is the one which contains one gram equivalent weight of the substance present in one litre of a solution.

22. Give few applications of Avogadro's hypothesis?

The main applications of Avogadro's hypothesis are

(i) Deducing the atomicity of elementary gases.

(ii) Deriving the relation between vapour density and relative molecular mass of a gas.

(iii) Mole concept.

(iv) Molar volume.

23. Give examples for gases whose atomicity is two?

The following gases will have atomicity two. (e.g.,) H_2, N_2, O_2, CI_2, F_2.

24. How will you find out the mass (a) One atom, (b) One molecule?

a) The mass of one atom = $\dfrac{\text{Gram atomic mass}}{\text{Avogadro number}}$

b) Mass of one molecule = $\dfrac{\text{Gram molecular mass}}{\text{Avogadro number}}$

25. What is the mass of 6.023 x 10^{23} molecules of ammonia?

6.023 x 10^{23} molecules of ammonia equal to one mole. Hence the mass is 17g.

26. How many molecules are present in 80 gm of sodium hydroxide?

The molecular mass of sodium hydroxide is = 40gm. = 6.023 x 10^{23} molecules

$\therefore$ 80 gm of sodium hydroxide is equal to = two mole.

i.e, 2 x 6.023 x 10^{23} molecules of sodium hydroxide.

27. Name the different methods of determining equivalent mass?

The equivalent mass can be determined by one of the following methods:

(i) Displacement method (ii) Oxide method (iii) Chloride method.

28. What is the principle involved in displacement method?

In displacement method hydrogen or oxygen or metal has been displaced from their compounds. From the amount of hydrogen or oxygen or metal displaced, the equivalent mass of an element is determined.

29. What is the principle involved in oxide method?

This method is used to find out the equivalent mass of those metals which can form their oxides. From the mass of metal and metallic oxide equivalent mass can be determined.

30. What is the principle involved in chloride method?

This method is used to find out the equivalent mass of those metals which can easily form their chlorides. From the mass of metal and metal chloride equivalent mass can be determined.

31. What is decinormal (0.1 N) solution?

A standard solution containing 1/10th of the gram equivalent of the substances is called decinormal (1/10 N or 0.1 N) solution.

32. Give the relationship between the mass of an atom and Avagadro number.

$\therefore$ The mass of one atom = $\dfrac{\text{Gram atomic mass}}{\text{Avogadro number}}$

33. Give the relationship between the mass of one molecule and Avogadro number.

$\therefore$ Mass of one molecule = $\dfrac{\text{Gram molecular mass}}{\text{Avogadro number}}$

34. Give the relationship between equivalent weight and valency.

$\therefore$ Equivalent weight = $\dfrac{\text{Atomic weight}}{\text{Valency}}$

35. Give the relationship between the molecular weight of the base and its equivalent mass.

$\therefore$ Equivalent mass of the base = $\dfrac{\text{Molecular mass of the base}}{\text{Acidity of the base}}$

36. Give relation between the equivalent mass of the salt and the valency of the metal?

$\therefore$ Equivalent mass of the salt = $\dfrac{\text{Molecular mass of the salt}}{\text{Valency of the metal in the salt}}$

37. Give the relationship between normality of the solution and the equivalent mass of the solute.

$\therefore$ Normality = Mass of solute present in one litre of the solution / ---

Equivalent weight of the solute

38. What is titration?

Titration is a process by which the volumes of solutions that react are determined experimentally.

39. Define end point.

The end or completion of a reaction detected experimentally by titration is known as the *end point*. If the reaction is an acid-base reaction, the end point is known as **neutralisation point**.

40. What is an indicator?

An indicator is a substance which, by a suitable colour change, indicates the completion of the reaction.

41. Name two metals whose equivalent weights are determined by oxide method.

The metals are copper, tin and iron.

III. ANSWER IN DETAIL:

1. How will you deduce the atomicity of elementary gases?

Definition: The number of atoms contained in one molecule of the element is called its atomicity.

Deduction of atomicity of Hydrogen: In the reaction between hydrogen and chlorine, 1 vol. of hydrogen combines with 1 vol. of chlorine to form 2 vols. of hydrogen chloride.

$$H_2 \quad + \quad Cl_2 \longrightarrow 2HCl$$

$$\underset{\text{(1 vol.)}}{\text{Hydrogen}} \quad + \quad \underset{\text{(1 vol.)}}{\text{Chlorine}} \longrightarrow \underset{\text{(2 vols.)}}{\text{Hydrogen chloride}}$$

Applying Avogadro's Law

$$\text{1 molecule} \quad + \quad \text{1 molecule} \longrightarrow \text{2 molecules}$$

From the above equation, we can say, each molecule of hydrogen chloride must contain atleast 1 atom of hydrogen.

Two molecules of hydrogen chloride, therefore, contain 2 atoms of hydrogen. These 2 atoms of hydrogen must have come from 1 molecule of hydrogen. Therefore, the molecule of hydrogen is **diatomic** and is written as H_2. Thus the atomicity of hydrogen is 2.

Similarly it can be shown that a molecule of chlorine contains 2 atoms, *i.e.,* the atomicity of chlorine is 2. From experimental evidence and similar arguments, it can be shown that oxygen and nitrogen also have 2 atoms per molecule, *i.e.,* their molecules are diatomic.

The molecules of hydrogen, chlorine, oxygen, and nitrogen are diatomic, *i.e.,* their atomicity is 2.

2. Derive the relationship between equivalent weight of atomic weight?

Consider an element X of atomic weight A and valency n, Suppose the element X combines with hydrogen to give the compound hydride of the formula XH_n

$$X + nH \longrightarrow XH_n$$

1 atom of X combines with n atoms of hydrogen or 1 gram-atom of X combines with *n* gram-atoms of hydrogen.

$$\text{Atomic weight of X} = A$$

$$\text{Atomic weight of hydrogen} = 1.008$$

$$n \times 1.008\text{g of hydrogen combines with} = A \text{ gram of X}$$

$$\therefore 1.008 \text{ grams of hydrogen combines with} = \frac{A}{n \times 1.008} \times 1.008$$

$$= \frac{A}{n}$$

By definition, this gives the equivalent weight of the element X.

$$\therefore \text{Equivalent weight} = \frac{\text{Atomic weight}}{\text{Valency}}$$

In the case of univalent elements, the atomic weight itself gives the equivalent weight. Elements having varying valency have different equivalent weight.

For example, Equivalent weight of copper in cuprous oxide

Cu_2O is 63.5. (here, valency of Cu is one).

Equivalent weight of copper in cupricoxide,

63.5

CuO is ------ = 31.75 (here, valency of Cu is 2).
2

3. How will you determine equivalent mass of an element by hydrogen displacement method?

Principle: This method is used to determine the equivalent mass of those metals (Sodium, Zinc, Aluminum, and Magnesium) which readily displace hydrogen from an acid or from a base or water.

$$Mg + 2\,HCl \longrightarrow MgCl_2 + H_2\uparrow$$

Procedure:

In this method a known mass of metal m *(g)* is added to acid or base or water. The volume of amount of hydrogen liberated (V_1) is measured at room temperature and pressure. The pressure of moist hydrogen gas displaced (P) is measured. From the aqueous tension (p), the pressure of dry hydrogen gas displaced (P_1) is measured.

Using the values of standard pressure (P_0) and standard temperature (T_0) and room temperature (T_1) the volume of hydrogen gas displaced is converted into volume of hydrogen gas displaced at STP by the gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_0V_0}{T_0}$$

Calculation:

The mass of the metal displaced by V_0 m^3 of the dry hydrogen gas at STP = m.g

The mass of the metal displaced by 1.12 x 10^{-2} m^{-3} = of the dry hydrogen gas at STP $= \dfrac{m}{V_0\,m^3} \times 1.12 \times 10^{-2} m^3$

The mass of 1.12 x 10^{-2} m^3 of at STP = 1.008g (hence by definition it represents the equivalent mass of the metal)

Equivalent mass of the metal $= \dfrac{\text{Mass of the metal x } 1.12 \text{ x } 10^{-2}\text{ m}}{V_0 m^3}$

4.How will you determine the equivalent weight of an element by chloride displacement method?

Principle: This method is used to find out the equivalent mass of those metals (Silver, Gold, Potassium, and Sodium) which can easily form their chlorides.

$$2Ag + Cl_2 \longrightarrow 2\,AgCl$$

Procedure:

In general a known mass of a metal (m_1) whose equivalent mass is to be determined is heated in the presence of chlorine. The mass of metal chloride (m_2), formed is found.

In some cases (Silver) the metals could not be directly converted into their chlorides. In such cases, known mass of metal (m_1) is converted into its salt by dissolving it in nitric acid. Then the metallic salt formed is converted into its chloride by adding dilute hydrochloric acid. The precipitated metal chloride is dried. The mass of metal chloride formed is found (m_2).

From the mass of metal chloride and the mass of metal, the mass of chlorine that has combined with the metal is found.

Calculation:

The mass of chlorine that has combined with the metal = $(m_2 - m_1)$g

$(m_2 - m_1)$ g of chlorine has combined with = m_1g of metal

∴ the 35.465g of chlorine combined with $= \left\{\dfrac{m_1}{m_2 - m_1}\right\}$ x 35.46

By definition it refers to the equivalent mass of the metal.

Equivalent mass of the metal $= \dfrac{\text{Mass of metal x 35.46}}{\text{Mass of chlorine that has combined with metal}}$

5. How will you find the equivalent weight of an acid? Explain with an example.

Equivalent mass of an acid: Equivalent mass of an acid is the number of parts by mass of the acids which contains 1.008 parts by mass of replaceable hydrogen.

Basicity of mineral acid is defined as the number of Replaceable hydrogen atoms present in one mole of the acid. Basicity of organic acid is defined as the number of carboxylic groups present in the acid.

$$\text{Equivalent mass of the Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the Acid}}$$

For monobasic acids the molecular mass and equivalent mass are the same.

Example: **Calculate the equivalent mass of hydrochloric acid.**

Solution:

The basicity of hydrochloric acid HCl is 1

The molecular mass of HCl = 1 + 35.46 = 36.46

$$\text{Equivalent mass of the Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the Acid}}$$

$$\text{Equivalent mass of HCl} = \frac{36.46}{1} = 36.46$$

6. How will you find the equivalent weight of an base? Explain with an example.

Equivalent mass of Bases: Equivalent mass of the base is the number of parts by mass of the base required to neutralize one equivalent mass of an acid.

Acidity of hydroxide base is defined as the number of replaceable hydroxyl ions present in one mole of the base.

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

For monoacidic bases the molecular mass and equivalent mass are the same.

Example: Calculate the equivalent mass of sodium hydroxide

Solution:

The acidity of sodium hydroxide NaOH is 1

The molecular mass of NaOH = (23 + 16 + 1) = 40

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

$$\text{Equivalent mass of NaOH} = \frac{40}{1} = 40$$

7. How will you find the equivalent weight of an salt? Explain with an example.

Equivalent mass of salts: Equivalent mass of a salt is the number of parts by mass of salt which reacts with one equivalent of mass of any other substance.

$$\text{Equivalent mass of the salt} = \frac{\text{Molecular mass of the salt}}{\text{Valency of the metal in the salt}}$$

Example: Calculate the equivalent mass of cupric chloride.

Solution:

The Valency of copper in cupric chloride $CuCl_2$ is 2

The molecular mass of $CuCl_2$ = 63.5 + (2 × 35.46) = 134.42

$$\text{Equivalent mass of the salt} = \frac{\text{Molecular mass of the salt}}{\text{Valency of the metal in the salt}}$$

$$\text{Equivalent mass of } CuCl_2 = \frac{134.42}{2} = 67.21$$

Note: In the case of salts like Na_2SO_4, the total valency of sodium is 2.

PROBLEMS

I. CALCULATION BASED ON V.D, GAW, GMW, M.W

1.Calculate the vapour density of a substance whose molecular mass is 96.

Solution:

$$\text{Vapour density} = \frac{\text{Molecular mass}}{2}$$

$$= \frac{96}{2} = 48$$

$\therefore$ The vapour density of a substance is 48.

2. Calculate the gram atomic weight of oxygen.

The atomic mass of oxygen = 16

$\therefore$ gram – atomic mass of oxygen (O) = 16 g

3. Calculate the gram molecular weight of hydrogen.

Molecular weight of hydrogen (H_2) = 2

$\therefore$ gram – molecular weight of hydrogen (H_2) = 2 g

4. Calculate the gram molecular weight of methane.

Molecular weight of methane (CH_4) = 16

$\therefore$ gram molecular weight of methane = 16 g

5. Calculate the molecular weight of water.

The Molecular formula of water is H_2O.

Atomic mass of H = 1

Atomic mass of O = 16

$\therefore$ Molecular mass of water

= (2 x atomic mass of H) + (1 x atomic mass of O) = 2 x 1 + 1 x 16 = 18

i.e., molecular mass of water = 18 amu.

6. Calculate the molecular weight of sulphuric acid.

The Molecular formula of sulphuric acid is H_2SO_4.

Atomic mass of H = 1

Atomic mass of O = 16

Atomic mass of S = 32

$\therefore$ Molecular mass of H_2SO_4

= (2 x atomic mass of H) + (1 x atomic mass of S) + (4 x atomic mass of O)

= (2 x 1) + (1 x 32) + (4 x 16) = 2 + 32 + 64 = 98

i.e., molecular mass of H_2SO_4 = 98 amu.

II. PROBLEMS BASED ON MOLE CONCEPT (T.B. PAGE 26 – 31)

Some useful formulae for problems on mole concept are:

1) (a) $\text{No. of moles} = \frac{\text{Mass}}{\text{Atomic mass}}$ (b) $\text{No. of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$

2) (a) $\text{No. of moles} = \frac{\text{No of atoms}}{6.023 \times 10^{23}}$ (b) $\text{No. of moles} = \frac{\text{No of molecules}}{6.023 \times 10^{23}}$

(i) Calculation of number of moles when the mass of the substance is given.

1. Calculate the number of moles in (a) 9.2 g sodium (b) 48g of oxygen atoms.

(i) Calculation of number of moles when the mass of the substance is given.

Solution: $\text{No. of moles} = \frac{\text{Mass}}{\text{Atomic mass}}$ for elements.

Atomic mass of sodium and oxygen are 23 and 16 respectively.

(a) $\therefore$ No. of moles in 9.2g sodium = 9.2/23 = 0.4

(b) $\therefore$ No. of moles of oxygen atoms = 48/16 = 3.0

2. Calculate the number moles in (a) 6g hydrogen molecules (b) 11 g carbon dioxide (c) 20g sodium hydroxide and (d) 24g water.

Solution: $\text{No. of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$

Molecular mass of H_2, CO_2, NaOH and water atoms are 2g, 44g, 40g and 18g respectively.

(a) $\therefore$ No. of moles in 6g Hydrogen molecules = 6 /2 = 3

(b) $\therefore$ No. of moles in 11g CO_2 molecules = 11/44 = 0.25

(c) $\therefore$ No. of moles in 20g NaOH molecules = 20/40 = 0.5

(d) $\therefore$ No. of moles in 24 g water = 24 /18 = 1.333

(ii). Calculation of mass when the number of moles are given:

Problem-1 Calculate the mass of the following in grams (a) 0.4 mole oxygen molecules (b) 2.5 mole carbon dioxide.

Solution: ***Mass = molecular mass*** **x** ***no. Of moles***

(a) Mass of 0.4 mole oxygen molecule = 32 x 0.4 = 12.8g

(b) Mass of 2.5 mole of CO_2 = 44 x 2.5 = 110g

Problem-2 Calculate the mass of the following in grams. (a) 0.25 mole of copper (b) 2.2 mole of nitrogen atoms.

Solution: ***Mass = atomic mass*** **x** ***No. of moles***

(a) Mass of 0.25 mole of copper = 63.54 x 0.25 = 15.89g

(b) Mass of 2.2 mole of nitrogen atoms = 14 x 2.2 = 30.8g

(iii) Calculation of the number of particles from the mass of the substance:

Problem-1 Calculate the numbers of molecules in (a) 66g carbon dioxide (b) 90g of glucose ($C_6H_{12}O_6$).

Solution: (a) Gram molecular mass of CO_2 = 44g

44g of CO_2 contains = 6.023×10^{23} molecules

The no. of molecules in 66g of CO_2 = $\frac{6.023 \times 10^{23} \times 66}{44} = 9.033 \times 10^{23}$

(b) Gram molecular mass of glucose ($C_6H_{12}O_6$) = 180.

Since 180g of glucose contains = 6.022×10^{23} molecules

The no. of molecules in 90g of it = $\frac{6.022 \times 10^{23} \times 90}{180} = 3.011 \times 10^{23}$

(iv) Calculate the mass from the number of particles of the substance.

Problem 1 Calculate the mass of ammonia containing 12.046×10^{23} molecules.

Solution: Gram molecular mass of NH_3 = 17

6.022×10^{23} molecules of ammonia weighs = 17 g

$\therefore$ The mass of 12.044×10^{23} molecules of ammonia is = $\frac{17}{6.022 \times 10^{23}} \times 12.044 \times 10^{23} = 34$ g

Problem 2 Calculate the mass of sucrose containing 1×10^{24} molecules.

Solution: Gram molecular mass of sucrose = 342

6.022 x 1023 molecules of sucrose weigh = 342 g

$\therefore$ The mass of 1 x 1024 molecules of sucrose is = $\frac{342}{6.022 \times 10^{23}} \times 1 \times 10^{24} = 560$ g

(v) Calculation of the number of moles when the number of atoms or molecules are given

Problem-1 Convert the following into number of moles (a) 1.8×10^{24} molecules of CO_2 (b) 1.204×10^{23} atoms of sodium.

Solution: (a) $\text{No. of moles} = \frac{\text{No. of molecules}}{\text{Avogadro number}} = \frac{1.8 \times 10^{24}}{6.023 \times 10^{23}} = 2.99$

(b) $\text{No. of moles} = \frac{\text{No. of atoms}}{\text{Avogadro number}} = \frac{1.204 \times 10^{23}}{6.023 \times 10^{23}} = 0.2$

(vi) Calculate the number of molecules when the number of moles are given

Problem –1 Calculate the number of molecules in 2.5 mole of water

Solution: No. of molecules = No. of moles x Avogadro number

$= 2.5 \times 6.023 \times 10^{23} = 1.5055 \times 10^{24}$

ADDITIONAL PROBLEMS BASED ON MOLE CONCEPT

1. Calculate the number of moles in 6 g of sodium. (Atomic mass of sodium = 23).

Solution:

$$\text{No of moles} = \frac{\text{Mass}}{\text{Atomic mass}}$$

$= 6 / 23 = 0.26$

$\therefore$ 6 g of sodium = 0.26 mole

2. How many grams of each of the following elements must be taken to get 1 mol of the element?

(a) Na (b) Cl (c) Cu

Solution:

The mass of 1 mol of an element is its atomic mass expressed in grams. The atomic masses of Na, Cl, Cu are 23, 35.5 and 63.5 g respectively. Therefore we must take,

(a) 23 g of Na (b) 35.5 g of Cl (c) 63.5 g of Cu

3. An atom of neon has a mass of 3.35×10^{-23} g. How many atoms of neon are there in 40 g of the gas?

Solution:

$$\text{Number of atoms} = \frac{\text{total mass}}{\text{Mass of 1 atom}} = \frac{40 \text{ g}}{3.35 \times 10\text{-}23 \text{ g}} = 11.94$$

4. Find the mass in grams of 3 mol of zinc.

Solution: The atomic mass of zinc = 65.41.

The mass of 1 mol = 65.41 grams of zinc

Thus, the mass of 3 mol of zinc = 65.41 x 3g

= 196.23 g.

5. How many atoms of copper are present in 0.5 mol of pure copper metal?

Solution: 1 mol of copper metal = 6.022×10^{23} atoms of Cu.

$\therefore$ 0.5 mol of copper metal = $0.5 \times 6.022 \times 10^{23}$ atoms of Cu.

= 3.011×10^{23} atoms of Cu.

6. What is the mass of 5 mol of ammonia?

Solution: Mass of ammonia = number of moles x gram –molecular mass

= 5 x 17 g

= 85 g.

7. What mass in grams is represented by (a) 0.40 mol of CO_2, (b) 3 mol of NH_3

(C = 12, O = 16, N = 14, H = 1).

Solution: Weight in grams = number of moles x molecular mass

Hence, (a) mass of CO_2 = 0.40 x 44 = 17.6 g

(b) mass of NH_3 = 3 x 17 = 51 g

8. Calculate the volume in litres of 20 g of hydrogen gas at STP.

Solution:

$$\text{Number of moles of hydrogen} = \frac{\text{Mass of hydrogen}}{\text{Molecular mass of hydrogen in grams}} = \frac{20 \text{ g}}{2 \text{ g}}$$

= 10 moles.

$\therefore$ Volume of hydrogen = Number of moles x standard molecular volume

= 10 x 22.4

= 224 litres.

9. The molecular mass of H_2SO_4 is 98 amu. Calculate the number of moles of each element in 294 g of H_2SO_4.

Solution: Number of moles of H_2SO_4 = 294 / 98 = 3

From the formula H_2SO_4 we know that 1 molecule of H_2SO_4 contains

2 atoms of H, 1 atom of S and 4 atoms of O.

Thus, 1 mol of H_2SO_4 will contain 2 mol of H

1 mol of S and 4 mol of O atoms.

Therefore, in 3 mol of H_2SO_4,

Number of moles of H = 2 x 3 = 6,
Number of moles of S = 1 x 3 = 3,
Number of moles of O = 4 x 3 = 12.

10. Find the mass of oxygen contained in 1 kg of potassium nitrate (KNO_3).

Solution: Since 1 molecule of KNO_3 contains 3 atoms of oxygen, 1 mol of KNO_3 contains 3 mol of oxygen atoms.

$\therefore$ moles of oxygen atoms = 3 x moles of KNO_3
= 3 x 1000 / 101
= 29.7
(molecular mass of KNO_3 = 101)

$\therefore$ mass of oxygen = number of moles x atomic mass = 29.7 x 16 g
= 475.2 g.

11.How many atoms are there in 10 grams of oxygen.

Solution: Molecular weight of oxygen = 16

16 grams of oxygen contains = 6.023×10^{23} atoms

$\therefore$ 10g of oxygen contains = $\dfrac{6.023 \times 10^{23}}{16} \times 10$

= 3.764×10^{23} atoms of oxygen

12.How many moles of hydrogen are there in 34 grams of ammonia.

Solution: Molecular mass of ammonia = 17 g

No. of moles = $\dfrac{\text{Mass}}{\text{Molecular mass}}$

= $\dfrac{34}{17}$ = 2 mole

34 grams of H_2 contains 2 moles of H_2.

13. How many molecules are there in 4 mole of CO_2.

Solution: No of molecules = No of moles $\times 6.023 \times 10^{23}$

= $4 \times 6.023 \times 10^{23}$

= 24×10^{23}

$\therefore$ 4 moles of CO_2 contain 24.092×10^{23} of CO_2 molecule.

14. How many atoms are there in 6 moles of oxygen

Solution: No of Atoms = No of moles $\times 6.023 \times 10^{23}$

= $6 \times 6.023 \times 10^{23}$

36.138 = 6×10^{23} atoms

15. Calculate the mass of 5 moles of CO_2

Solution:

Molecular mass of CO_2 = 12 + 2 (16)

= 12 + 32 = 44 g = 1 mole

Mass of 5 moles of CO_2 = 44×5

= 220 g

16. Calculate the number of moles in 5.3 g of Na_2CO_3.

Solution:

Molar mass of Na_2CO_3 = $(2 \times 23) + (1 \times 12) + (3 \times 16)$

= 46 + 12 + 48 = 106

No of moles = =

= 0.05 mol

$\therefore$ 5.3 g of Na_2CO_3 contains 0.05 mole

17. Calculate the no of NH_3 molecules in 7 mole of NH_3 and mass of 7 mole of NH_3.

Solution:

Molecular mass of NH_3 = $(14 \times 1) + (1 \times 3)$

= $14 + 3 = 17$

(i) No of NH_3 molecules in 7 mole

No of Molecules = No of moles $\times$ 6.023×10^{23}

= $7 \times 6.023 \times 10^{23}$

= 42.161×10^{23}

$\therefore$ 7 mole of NH_3 contains 30.1×10^{23} molecules of NH_3.

(ii) Mass of 7 mole of NH_3 = No of moles $\times$ Molecular mass

= 7×17

= 119 g

7 moles of NH_3 contain 85.0 g of NH_3.

18.How many atoms are there in 8 g of oxygen.

No of oxygen atoms = No of moles $\times 6.023 \times 10^{23}$

$$\text{No of moles} = \frac{\text{Mass}}{\text{Atomic mass}} = \frac{8}{16} = 0.5 \text{ mole}$$

$\therefore$ No of oxygen atoms = $0.5 \text{ x } 6.023 \text{ x } 10^{23} = 3.0115 \text{ x}10^{23}$

$\therefore$ No of oxygen atoms present in 8 g of oxygen = $3.0115 \text{ x}10^{23}$

19. Calculate the mass of an atom of iron (Atomic mass of iron is 55.9).

Solution: Gram atomic mass of iron = 55.9 g

$$\text{Mass of one atom} = \frac{\text{Gram atomic mass}}{\text{Avogadro number}} = \frac{55.9}{6.023 \text{ x } 10^{23}} = 9.283 \text{ x } 10^{-23} \text{ g}$$

20. Mass of an atom of an element is $6.5 \text{ x } 10^{-23}$ g. Calculate its atomic mass.

Solution: Gram atomic mass = mass of 1 atom x Avogadro number

= $6.5 \text{ x } 10^{-23} \text{ x } 6.023 \text{ x } 10^{23} = 39.143$ g

$\therefore$ Atomic mass = 39.143

One gram molecular mass of any substance contains Avogadro number of molecules.

$$\therefore \text{Mass of one molecule} = \frac{\text{Gram molecular mass}}{\text{Avogadro number}}$$

21. Calculate the mass of one molecule of CO_2.

Solution: Gram molecular mass of CO_2 = 44g

$$\therefore \text{Mass of one molecule} = \frac{\text{Gram molecular mass}}{\text{Avogadro number}} = \frac{44}{6.023 \text{ x } 10^{23}} = 7.306 \text{ x } 10^{-23} \text{ g}$$

22. Mass of one molecule of a substance is $3.4 \text{ x } 10^{-22}$ g. Calculate its molecular mass.

Solution: Gram molecular mass = Mass of 1 molecule x Avogadro number

= $3.4 \text{ x } 10^{-22} \text{ x } 6.023 \text{ x } 10^{23} = 204.7$ g.

Note: For any substance, the mass of one mole of the substance = its gram molecular mass.

Eg. : (1) Mass of one mole ($6.023\text{x}10^{23}$ molcules) of water = 18g

(2) Mass of one mole ($6.023\text{x}10^{23}$ molcules) of ozone = 48g

(3) Mass of one mole ($6.023\text{x}10^{23}$ molcules) of CO_2 = 44g

23. Convert the following quantities into moles (i) 24.4 g oxygen gas, (ii) 10 g methane, (iii) 25.3 g sodium, (iv) 15.0 g acetic acid.

(at wts. H = 1, C = 12, O = 16, Na = 23).

$$\text{No.of moles} = \frac{\text{Weight of the substance in g (w)}}{} = \frac{w}{}$$

Molecular weight of the substance (m) M

(i) Molecular weight of oxygen is 32.

$\therefore$ No. of moles of oxygen $= \dfrac{24.4}{32} =$ **0.7625 mole.**

(ii) Molecular formula of methane is CH_4

$\therefore$ Its molecular weight $= 12 + 4 \times 1 = 16$

$\therefore$ No. of moles of methane $= \dfrac{10}{16} =$ **0.625 mole.**

(iii) For sodium, its atomic weight and molecular weight are the same, namely 23.

$\therefore$ No. of moles of sodium $= \dfrac{25.3}{25} =$ **1.1 mole.**

(iv) Molecular formula of acetic acid is CH_3 COOH i.e. $C_2H_4O_2$.

Its molecular weight $=(12 \times 2) + (4 \times 1) + (16 \times 2) = 60$

$\therefore$ No. of moles of acetic acid $= \dfrac{15}{60} =$ **0.25 mole.**

24. Convert the number of moles of the following substance into their respective weights (i) 0.21 mole chlorine gas, (ii) 0.81 mole water, (iii) 0.6 mole oxygen atoms, (iv) 0.3714 mole calcium carbonate, (v) 3.7 moles helium, (vi) 0.1 mole OH^- ions.

(At. Wts. H = 1, He = 4, C = 12, O = 16, Cl = 35.5, Ca = 40).

Weight of the substance (w g) = moles (n) x molecuar weight (M)

w = n x M.

(i) Molecular weight of chlorine is 71.

$\therefore$ wt. of chlorine $= 0.21 \times 71 =$ **14.91 g**

(ii) Molecular weight of H_2O $= (2 \times 1) + 16 = 18.$

$\therefore$ wt. of water $= 0.81 \times 18 =$ **14.58 g**

(iii) Atomic weight of oxygen $= 16$

$\therefore$ wt. of oxygen $= 0.6 \times 16 =$ **9.6 g**

(iv) Molecular formula of calcium carbonate is $CaCO_3$.

$\therefore$ Its molecular weight $= (40 \times 1) + (12 \times 1) + (16 \times 3) = 100$

$\therefore$ wt. of calcium carbonate $= 0.3174 \times 100 =$ **31.74 g**

(v) Atomic weight of helium and its molecular weight are the same, namely 4.

$\therefore$ wt. of helium $= 3.7 \times 4 =$ **14.8 g**

(vi) One mole OH^- ions is the same as its one gram ion namely 1 + 16 = 17 g

$\therefore$ wt. of OH^- ions $= 0.1 \times 17 =$ **1.7 g**

25. Calculate gram equivalents of the following: (i) 4 g oxygen, (ii) 7 g chlorine, (iii) 4.5 g aluminium, (iv) 2.4 g magnesium.

(Given: Equivalent weights, O = 8, Cl = 35.5, Al = 9, Mg = 12).

$$\text{Gram equivalent} = \frac{\text{weight of the element in g (w)}}{\text{Equivalent weight of the element (E)}} = \frac{W}{E}$$

$\therefore$ (i) gram equivalent of oxygen $= \dfrac{4}{8} =$ **0.5**

(ii) gram equivalent of chlorine $= \dfrac{7}{35.5} =$ **0.1972**

(iii) gram equivalent of aluminium $= \dfrac{4.5}{9} =$ **0.5**

(iv) gram equivalent of magnesium $= \dfrac{2.4}{12} =$ **0.2**

A. HYDROGEN DISPLACEMENT METHOD:

1. 0.180 g of a metal evolved 1.066×10^{-4} m^3 of hydrogen collected over water at a temperature of 290 K and a pressure of 1.0329×10^5 Nm^{-2}. The aqueous tension at290 K = 1.9193×10^3 Nm^{-2}. Calculate the equivalent mass of the metal.

Solution:

Lab values	STP values
$V_1 = 1.066 \times 10^{-4}$ m^3 $P_1 = (P - p$ Nm$^{-2})$ $= (1.0329 \times 10^5 - 1.9193 \times 10^3)$ $= 1.0137 \times 10^5$ Nm^{-2} $T_1 = 290$ K	$V_0 = ?$ $P_0 = 1.013 \times 10^5$ Nm^{-2} $T_0 = 273$ K

$$\frac{P_1V_1}{T_1} = \frac{P_0V_0}{T_0}$$

i.e., $P_1V_1T_0 = P_0V_0T_1$

$$V_0 = \frac{P_1V_1T_0}{T_1P_0} = \frac{1.0137 \times 10^5 \times 1.066 \times 10^{-4} \times 273}{290 \times 1.013 \times 10^5}$$

$V_0 = 1.0042 \times 10^{-4}$ m^3.

= Volume of hydrogen at STP.

At STP, 1.0042×10^{-4} m^3 of hydrogen is displaced by 0.206 g of metal.

At STP, 1.12 x 10-2 m3 of hydrogen will be displaced by, $= \dfrac{1.12 \times 10^{-2} \times 0.180}{1.0042 \times 10^{-4}}$

The equivalent mass of the metal = 20.

2. 1.420 g of a metal on reacting with dilute sulphuric acid gave 512.6 cc of hydrogen collected over water at 27°C and 756.5 mm pressure. Aqueous tension at 27°C is 26.5 mm. Calculate the equivalent weight of metal.

Solution:

Lab values	STP values
$V_1 = 512.6$ cc $P_1 = (P - p$ Nm$^{-2})$ $= (756.5 - 26.5)$ $= 730$ mm $T_1 = 300$ K	$V_0 = ?$ $P_0 = 760$ mm $T_0 = 273$ K

We know $$\frac{P_0V_0}{T_0} = \frac{P_1V_1}{T_1}$$

i.e., $$V_0 = \frac{P_1V_1}{T_1} \times \frac{T_0}{P_0}$$

$$= \frac{730 \times 512.6}{300} \times \frac{273}{760} = 448\text{cc.}$$

301

$\therefore$ Mass of hydrogen at STP = Volume of hydrogen x density of hydrogen

= 448 x 0.00009

= 0.04033 g.

Equivalent weight of the metal $= \dfrac{\text{Given mass of the metal}}{\text{Mass of hydrogen at STP}} \times 1.008$

1.420

$$= \frac{\text{-----------}}{0.04033} \times 1.008 = 35.5$$

$\therefore$ Equivalent weight of the metal = 35.5

3. 0.378 g of a metal reacts with dilute HCl completely and displaces hydrogen the volume of which is 2.181×10^{-4} m^3 at a temperature of 300 K and pressure of 1.041×10^5 Nm^{-2}. Calculate the equivalent mass of the metal. The value of aqueous tension at 300 K is 3.561×10^3 Nm^{-2}.

Solution:

Lab values	STP values
$V_1 = 0.378 \times 10^{-4}$ m^3 $P_1 = (P - p$ Nm$^{-2})$ $= (1.041 \times 10^5 - 3.561 \times 10^3)$ $= 1.0054 \times 10^5$ Nm^{-2} $T_1 = 300$ K	$V_0 = ?$ $P_0 = 1.013 \times 10^5$ Nm^{-2} $T_0 = 273$ K

We know $= \dfrac{P_0V_0}{T_0} = \dfrac{P_1V_1}{T_1}$

$$\therefore V_0 = \frac{1.0054 \times 10^5 \times 2.181 \times 10^{-4}}{\substack{300 \\ 301}} = \frac{273}{1.013 \times 10^5}$$

$= 1.970 \times 10\text{-}4$ m^3.

At STP 1.970×10^{-4} m^3 of hydrogen is displaced by 0.378 g of metal

At STP the mass of the metal which will displace 1.12×10^{-2} m^3 of hydrogen.

$$= \frac{378}{1.970 \times 10^{-4}} \times 1.12 \times 10^{-2} \text{ of the metal}$$

$\therefore$ The equivalent mass of the metal = 21.5

4. 0.40 gm of a metal, when dissolved in dil. HCl, displacement 2.95×10^{-4} m^3 of hydrogen collected over water at a temperature of 290 K and a pressure of 9.86×10^4 Nm^{-2}. Calculate the equivalent weight of the metal. Aqueous tension at 290 K = 1.92×10^3 Nm^{-2}.

Solution:

Lab values	STP values
$V_1 = 2.95 \times 10^{-4}$ m^3 $P_1 = (P - p$ Nm$^{-2})$ $= (9.86 \times 10^5 - 1.92 \times 10^3)$ $= 9.8408 \times 10^4$ Nm^{-2} $T_1 = 290$ K	$V_0 = ?$ $P_0 = 1.013 \times 10^5$ Nm^{-2} $T_0 = 273$ K

$$\frac{P_0 V_0}{T_0} = \frac{P_1 V_1}{T_1}$$

$$V_0 = 2.95 \times 10^{-4} \times \frac{273}{290} \times \frac{9.8408 \times 10^4}{1.013 \times 10^5} \text{ m}^3$$

$= 2.6977 \times 10^{-4}$ m^3

2.652×10^{-4} m^3 of hydrogen at STP is displaced by 0.40 gm of the metal.

1.12×10^{-2} m^3 of hydrogen at STP is displaced by $\dfrac{0.40 \times 1.12 \times 10^{-2}}{2.6977 \times 10^{-4}}$

$= 16.60$

This must be the gram equivalent weight of the metal.

Thus the equivalent weight of the metal = 16.60

5. In the determination of the equivalent weight of a metal by the hydrogen displacement method, 0.2020 g of a metal liberated 3.7×10^{-5} m^3 of dry hydrogen at a temperature of 285 K and a pressure of 1.04×10^5 Nm^{-2}. Calculate the equivalent weight of the metal.

Solution:

Lab values	STP values
$V_1 = 3.7 \times 10^{-5}\ m^3$ $P_1 = 1.04 \times 10^5\ Nm^{-2}$ $T_1 = 300$ K	$V_0 = ?$ $P_0 = 1.013 \times 10^5\ Nm^{-2}$ $T_0 = 273$ K

The volume of dry hydrogen is to be reduced to STP $\frac{P_0V_0}{T_0} = \frac{PV}{T}$

$$V0 = 3.7 \times 10^{-5} \times \frac{273}{300} \times \frac{1.04 \times 10^5}{1.013 \times 10^5} = 3.456 \times 10^{-5}\ m^3$$

3.456×10^{-5} m^3 of hydrogen at STP is displaced by 0.1063 g of metal.

1.12×10^{-2} m^3 of hydrogen at STP is displaced by $\frac{0.1063 \times 1.12 \times 10^{-2}}{3.456 \times 10^{-5}} = 34.44$

∴ The equivalent weight of the metal = 34.44

6. 1.308 g of a metal on reacting with dilute sulphuric acid gave 5.126×10^{-4} m^3 of hydrogen collected over water at 270°C and 1.008×10^5 Nm^{-2}. Aqueous tension at 27°C is 3.5322×10^3. Calculate the equivalent mass of the metal. (T. B. Page 34)

The pressure of Dry hydrogen displaced (P_1) = The Pressure of moist hydrogen displaced (P) – Aqueous tension (p)

$= 1.008 \times 10^5 - 3.5322 \times 10^3$
$= 100.8 \times 10^3 - 3.5322 \times 10^3$
$= (100.8 - 3.5322) \times 10^3$
$= 97.2678 \times 10^3$
$= 9.72678 \times 10^4\ Nm^{-2}$

The pressure of dry hydrogen displaced (P_1) $= 9.72678 \times 10^4\ Nm^{-2}$
The volume of hydrogen displaced (V_1) $= 5.126 \times 10^{-4}\ m^3$
The room temperature (T_1) = 270°C + 273 = 300 K
Standard Pressure (P_0) $= 1.013 \times 10^5$
Standard Temperature (T_0) = 273 K

$$\frac{P_1V_1}{T_1} = \frac{P_0 V_0}{T_0}$$

$$V_0 = P_1V_1 T_0 / P_0 T_1$$

$= (9.72678 \times 10^4 \times 5.126 \times 10^{-4} \times 273) / (1.013 \times 10^5 \times 300)$

$= 4.479\ \ 10^{-4}\ m^3$

At STP $4.479\ \ 10^{-4}$ m^3 of hydrogen is displaced by 1.308 g of metal

Therefore at STP 1.12×10^{-2} m^3 of hydrogen will be Displaced by $\Big\}$ $= (1.308 \times 1.12 \times 10^{-2}) / 4.479\ \ 10^{-4}$

Equivalent mass of the metal = 32.7

B. PROBLEMS BASED ON OXIDE METHOD:

1. In an experiment, 4.02 g of zinc gives 4.42 g of zinc oxide. Calculate the equivalent weight of zinc.

Solution: Weight of zinc oxide = 4.42 g

Weight of zinc = 4.02 g

Weight of oxygen = 0.40 g

By definition, equivalent mass of the metal $= \dfrac{\text{Mass of the metal}}{\text{Mass of oxygen}} \times 8$

$= \dfrac{4.02}{0.40} \times 8 = 80.4$ g

Thus the equivalent weight of the metal is 80.4.

2. 0.44 g of a metal gives on oxidation 0.60 g of its oxide. Calculate the equivalent mass of the metal.

Solution: Mass of the metal oxide = 0.60

Mass of metal = 0.44

Mass of oxygen = 0.16

Equivalent mass of the metal $= \dfrac{\text{Mass of the metal}}{\text{Mass of oxygen}} \times 8$

$\dfrac{0.44}{0.16} \times 8 = 22$

Equivalent mass of the metal = 22.

3. On heating 2.80 g of a metal oxide, in a current of dry hydrogen 2.30 g of the metal are obtained after the completion of the reaction. Calculate the equivalent weight of the metal.

Solution: Mass of the metal oxide = 2.80 g

Mass of the metal obtained = 2.30 g

Weight of oxygen = 0.50 g

Equivalent weight of the metal $= \dfrac{\text{Mass of the metal}}{\text{Mass of oxygen}} \times 8$

$= \dfrac{2.30}{0.50} \times 8 = 36.8$

Equivalent weight of the metal = 36.8.

4. On heating 0.9420 g of Cu (II) oxide in a current of hydrogen, the resulting copper weighed 0.7715 g. What is the equivalent weight of copper?

Solution: Weight of cupric oxide = 0.9420 g

Weight of copper = 0.7715 g

Weight of oxygen = 0.1705 g

By definition,

Equivalent mass of metal $= \dfrac{\text{Mass of the metal}}{\text{Mass of oxygen}} \times 8$

$= \dfrac{0.7715}{0.1705} \times 8 = 36.2$

Thus the equivalent weight of copper = 36.2.

C. PROBLEMS BASED ON CHLORIDE:

1. The chloride of a metal contained 52.20% of the metal. What is the equivalent weight of the metal?

Solution: Weight of metal chloride = 100 g

Weight of metal = 52.20g

∴ Weight of chlorine = 100 – 52.20

= 47.80

47.15 g of chloride combines with 47.80 g of metal.

∴ 35.46 g of chlorine will combine with

$$= \frac{52.20}{47.80} \times 35.46 = 39$$

Equivalent weight of metal = 39.

2. The chloride of a metal contains 30% of chlorine. Calculate the equivalent mass of the metal.

Solution: Mass of the metal chloride = 100 g

Mass of chlorine = 30 g

∴ Mass of the metal = 100 – 30 = 70 g

By definition,

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of the metal}}{\text{Mass of chlorine}} \times 35.46$$

$$= \frac{75}{30} \times 35.46 = 88.65$$

∴ The equivalent mass of the metal = 88.65.

3. 0.532 g of a metal gives 0.666 g of its chloride. Calculate the equivalent mass of the metal. Solution:

Mass of the metal chloride = 0.666g

Mass of metal = 0.532g

Mass of chlorine = 0.134g

By definition,

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of the metal}}{\text{Mass of chlorine}} \times 35.46$$

$$= \frac{0.532}{0.134} \times 35.46 = 140.78$$

∴ The equivalent mass of the metal = 140.78.

4. 0.555 g of a metal gives 0.444 g of its chloride. Calculate the equivalent mass of the metal.

Solution: Mass of the metal chloride = 0.555g

Mass of metal = 0.444g

Mass of chlorine = 0.111g

By definition,

$$\text{Equivalent mass of the metal} = \frac{\text{Mass of the metal}}{\text{Mass of chlorine}} \times 35.46$$

$$= \frac{0.444}{0.111} \times 35.46 = 141.84$$

∴ The equivalent mass of the metal = 141.84.

Equivalent weight of the metal = 15.21

V. CALCULATION OF EQUIVALENT MASS OF COMPOUNDS:

1. An acid has molecular weight 98 and its equivalent weights is 49. Find its basicity.

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

Basicity = 98 / 49 = 2.

2. A base has molecular weight 78. It has three hydroxyl ions. What is its equivalent weight?

$$\text{Equivalent weight of base} = \frac{\text{Molecular weight}}{\text{Number of hydroxyl ions}}$$

Equivalent weight of base = 78 / 3 = 26.

3. Calculate the equivalent mass of hydrochloric acid.

The molecular mass of HCl = 1 + 35.46 = 36.46

The basicity of hydrochloric acid HCl is 1.

$$\text{Equivalent mass of the Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the Acid}}$$

Equivalent mass of HCl = 36.46 / 1 = 36.46

4. Calculate the equivalent mass of cupric chloride.

The molecular mass of $CuCl_2$ = 63.5 + (2 x35.46) = 134.42

The valency of copper in cupric chloride $CuCl_2$ is 2.

$$\text{The molecular mass of the salt} = \frac{\text{Molecular mass of the salt}}{\text{Valency of the metal in the salt}}$$

Equivalent mass of $CuCl_2$ = 134.42 / 2 = 67.21

5. Calculate the equivalent mass of Sulphuric acid.

The molecular mass of H_2SO_4 = (2 x1) + (1 x32) + (4 x 16) = 98

The basiscity of sulphuric acid H_2SO_4 is 2.

$$\text{Equivalent mass of the Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the Acid}}$$

Equivalent mass of H_2SO_4 = 98 / 2 = 49.

6. Calculate the equivalent mass of sodium hydroxide.

The molecular mass of NaOH = (23 + 16 + 1) = 40

The acidity of sodium hydroxide NaOH is 1.

$$\text{Equivalent mass of the base} = \frac{\text{Molecular mass of the Base}}{\text{Acidity of the Base}}$$

Equivalent mass of NaOH = 40 / 1 = 40.

7. Calculate the equivalent mass of Acetic acid.

The molecular mass of CH_3COOH = (2 x 12) + (4 x 1) + (2 x 16) = 60

The basicity of acetic acid CH_3COOH is 1.

$$\text{Equivalent mass of Acid} = \frac{\text{Molecular mass of the Acid}}{\text{Basicity of the acid}}$$

Equivalent mass of Acetic acid = 60 / 1 = 60.

8. Calculate the equivalent mass of ortho phosphoric acid?

Solution:

The molecular mass of H_3PO_4 = (1 x 3) + (1 x 31) + (4 x 16)
= 3 + 31 + 64 = 98.

The basicity of ortho phosphoric acid (H_3PO_4) = 3

$$\text{Equivalent mass of } H_3PO_4 = \frac{\text{Molecular mass}}{\text{Basicity}}$$

$$\frac{98}{3} = 32.67$$

9. Calculate the equivalent mass of $Al(OH)_3$.

Solution:

The molecular mass of Al $(OH)_3$ = (1 x 27) + (3 x16) + (3 x 1)
= 27 + 48 + 3 = 78

The acidity of aluminium hydroxide is 3

$$\text{Equivalent mass of the base} = \frac{\text{Molecular mass}}{\text{Acidity of the base}}$$

$$= \frac{78}{3} = 26$$

10. Calculate the equivalent mass of calcium chloride ($CaCl_2$).

Solution:

The molecular mass of $CaCl_2$ = 40 + (35.4 x 2)
= 40 + 71 = 111.

The valency of calcium in calcium chloride is 2.

$$\text{Equivalent mass of } CaCl_2 = \frac{\text{Molecular mass of the salt}}{\text{Valency of the metal}}$$

$$= \frac{111}{2} = 55.5$$

11. Calculate the equivalent mass of Sodium sulphate (Na_2SO_4).

Solution: Na_2SO_4 has 2 atoms of sodium in one molecule and so the total valency of sodium is 2 x 1 = 2.

The molecular weight of Na_2SO_4 = (2 x 23) + 32 + (4 x 16)
= 46 + 32 + 64 = 142.

$$\text{Equivalent weight of } Na_2SO_4 = \frac{\text{Molecular mass of the salt}}{\text{Total valency of the metal}}$$

$$= \frac{142}{2} = 71.$$

VI. CALCULATION BASED ON VOLUMETRIC ANALYSIS:(T.B. PAGE 42 – 44)

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Example 16

80 g of NaOH pellets dissolved in 1 litre water. What is the normality?

Equivalent mass of the sodium hydroxide = 40

$$\text{Normality} = \frac{\text{Mass in grams per litre of solution}}{\text{Equivalent mass}} = 80/40 = 2$$

The normality of sodium hydroxide solution containing 80 gram of sodium hydroxide is 2.

Example 17

How much 12 N HCl do you need to make 400 ml of 2N solution?

$$V_1N_1 = V_2N_2$$

$$V_1 \text{ X } (12N) = (400 \text{ ml}) \text{ x } (2N)$$

$$V_1 = 400 \text{ x } 2 \text{ ml} / 12 = 66.67 \text{ ml}$$

∴66.7 ml of 12 N HCl is needed to prepare 400 ml of 2N HCl solution.

Example 18

Take 100 ml of a concentrated acid and make it 2 liters of 0.5N solution by adding water. What is the normality of the solution?

$$V_1N_1 = V_2N_2$$

$$(100 \text{ ml})\ N_1 = (2L)\ (0.5N)$$

Units on both sides of the formula must agree. Since the units for volume are not the same, 2L should be converted to 2,000 ml (or 100 ml converted to 0.1 L)

$$(100 \text{ ml})\ N_1 = (2{,}000 \text{ ml}) \text{ x } (0.5N)$$

$$N_1 = (2000 \text{ ml} \times 0.5 \text{ N}) / 100 \text{ ml}$$

$$N_1 = 10 \text{ N}$$

Example 19

10 ml of HCl is required to neutralize 15 ml of 1N NaOH. What is the concentration of HCl?

$$V_1N_1 = V_2N_2$$

$$(10 \text{ ml })\text{x } (N_1) = (15 \text{ ml}) \text{ x } (1\text{N})$$

$$N_1 = 1.5\text{N}$$

Example 20

What volume of 0.1 N HNO_3 is required to neutralize completely 50 ml of a 0.15 N solution of $Ba(OH)_2$?

$$V_1N_1 = V_2N_2$$

$$V_1 \text{ X } 0.1 \text{ N} = 50 \text{ ml X } 0.15 \text{ N}$$

$$V_1 = (50 \text{ X } 0.15) / 0.1$$

$$V_1 = 75 \text{ ml}$$

Volume of HNO_3 = 75 Ml

ADDITIONAL PROBLEMS

1. 40 g of NaOH pellets dissolved in 1 litre water. What is the normality?

Equivalent mass of the sodium hydroxide = 40

$$\text{Normality} = \frac{\text{Mass in grams per litre of solution}}{\text{Equivalent mass}}$$

= 40 / 40 = 1.

2. How much 6 N HCl do you need to make 200 ml of 2N solution?

$$V_1N_1 = V_2N_2$$

$$V_1 \text{ x } (6\text{N}) = (200 \text{ ml}) \text{ x } (2\text{N})$$

$$V_1 = 400 / 6 = 66.66 \text{ ml}$$

∴66.66 ml of HCl is required to make 200 ml of 2N acid solution.

3. Take 50 ml of a concentrated acid and make it 2 litres of 1N solution by adding water. What is the normality of the solution?

$$V_1N_1 = V_2 N_2$$

$$(50 \text{ ml}) N_1 = (2\text{L}) (1 \text{ N})$$

$$(100 \text{ ml}) N_1 = (2{,}000 \text{ ml}) \text{ x } (1\text{N})$$

$$N_1 = \frac{(2000 \text{ ml x } 1\text{N})}{100 \text{ ml}}$$

$$N_1 = 20 \text{ N}$$

4. 20 ml of HCl is required to neutralize 25 ml of 1N NaOH. What is the concentration of HCl?

$$V_1N_1 = V_2 N_2$$

$$(20 \text{ ml}) \text{ x } (N_1) = 25 \text{ ml}) \text{ x } (1\text{N})$$

$$N_1 = 25 / 20 = 1.25 \text{ N}$$

5. How much 12N HCl do you need to make 500 ml of 2.5 N

Solution:

$$V_1 N_1 = V_2 N_2$$

$$V_1 \text{ x } 2\text{N}= 500 \text{ x } 2.5$$

$$V_1 = \frac{1250}{2} = 625 \text{ ml}$$

833.33 ml of 1.5 N HCL is required to make 500 ml of 2.5 N solution.

6. 25 ml of HCl is required to neutralize 15 ml of 0.5 N NaOH. What is the concentration of HCl?

Solution: $V_1 N_1 = V_2 N_2$

$30 \times N_1 = 15 \times 0.5\ N$

$$N_1 = \frac{30 \times 0.5}{25} = 0.6\ N$$

Normality of HCl = 0.3 N

7. What volume of 0.5 N HNO_3 is required to neutralize completely 40 ml of 0.1 N solution of Barium hydroxide?

Solution: $V_1 N_1 = V_2 N_2$

$V_1 \times 0.5 = 40 \times 0.1$

$$V_1 = \frac{40 \times 0.1}{0.5} = 8\ ml$$

Volume of HNO_3 required = 8 ml

8. What should be the normality of a solution prepared by diluting 500 ml of 0.400 N sulphuric acid with 1000 ml of water?

Solution: $V_1 N_1 = V_2 N_2$

(Original solution) (Final solution)

$0.400 \times 500 = 1500 \times N_2$

$$N_2 = \frac{500 \times 0.4}{1500}$$

$= 0.133$ N

9. 40 ml of 0.2 N oxalic acid solution is titrated against 0.1 N solution of NaOH. What volume of NaOH will be required for complete neutralization?

Solution: Volume of oxalic acid V_1 = 40

Normality of oxalic acid N_1 = 0.2 N

Volume of NaOH V_2 = ?

Normality of NaOH N_2 = 0.1 N

$V_1 N_1 = V_2 N_2$

$40 \times 0.2 = V_2 \times 0.1$

$$V_2 = \frac{40 \times 0.2}{0.1}$$

= 80 ml.

10. 20 ml of 0.1N NaOH solution neutralizes 25 ml of a solution of nitric acid. Calculate the strength of the nitric acid solution?

Solution: Volume of NaOH solution = V_1 = 20 ml

Normality of NaOH solution = N_1 = 0.1N

Volume of HNO_3 solution = V_2 = 25ml

Normality of HNO_3 solution = N_2 = ?

$V_1 N_1 = V_2 N_2$

$20 \times 0.1 = 25 \times N_2$

$$N_2 = \frac{20 \times 0.1}{25}$$

$\therefore$ The strength of nitric acid solution = 0.08 N.

3. LAWS OF CHEMICAL COMBINATIONS

Introduction:

Compounds are formed by chemical combination of reactants (atoms or molecules) which may be solid, liquid or gaseous. Chemical combination occurs in definite proportion by weight or by volume. Based on various experiments performed by different scientists, the laws of chemical combinations were formulated.

These laws laid the foundation of *stoichiometry*, a branch of chemistry in which quantitative relationship between masses of reactants and products is established. The study of these laws led to the development of a theory concerning the nature of matter.

There are five laws of chemical combinations. The first four deal with combination of substances by weight and

the fifth with combination of gases by volume.

IMPORTANT TERMS & DEFINITIONS

Law of mass action: "The total mass of substances taking part in a chemical reaction remains the same throughout the change."

Law of multiple proportion: "When two elements A and B combine to form two or more compounds, then different weights of B which combine with a fixed weight of A bears a simple numerical ratio to one another".

Law of conservation of mass: When two elements combine separately with a definite mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.

Limitations of Law multiple proportion: The law is valid till an element is present in one particular isotopic form in all its compounds. When an element exists in the form of different isotopes in its compounds, the law does not hold good.

Stoichiometry a branch of chemistry in which quantitative relationship between masses of reactants and products are established. The study of these laws led to the development of a theory concerning the nature of matter.

SELF EVALUATION (T.B.PAGE 60)

I. Choose the correct answer.

1. .... is a branch of chemistry in which quantitative relationship between masses of reactants and products are established.
   (a) Stoichiometry (b) Physical chemistry
   (c) Organic chemistry (d) Quantitative chemistry
2. Law of conservation of mass was established by ....
   (a) Lavoisier (b) Lomonssoff (c) Dalton (d) Berzelius
3. Law of multiple proportion was enunciated by ....
   (a) Dalton (b) Lavoisier (c) Berzelius (d) Proust
4. Law of combining volume was given by ....
   (a) Gay lussac (b) Lavoisier (c) Berzelius (d) Proust
5. Some basic experiments on the law of reciprocal proportions was done by ....
   (a) Wenzel (b) Lavoisier (c) Dalton (d) Gay lussac
6. The ratio of the volumes of reactants and products in the formation of NH_3 is ....
   (a) 1 : 3 : 2 (b) 1 : 2 : 3 (c) 1 : 1 : 2 (d) None
7. When two or more gases react with one another, their volumes bear simple ratio. This statement given by ....
   (a) Law of mass action (b) Laws of multiple proportion
   (c) Law of reciprocal proportion (d) Law of combining volume
8. Law of definite proportion was stated in the year ....
   (a) 1756 (b) 1774 (c) 1799 (d) 1803
9. Experimental verification for the law of multiple proportion was verified by ....
   (a) Berzelius (b) Dalton (c) Lavoisier (d) lomonssoff
10. Law of Multiple proportion will not hold good for elements with different .... in its compounds.
    (a) Isotopes (b)Isomers (c) isobars (d) Vapour pressure

Answers:

1. (a) 2. (a) 3. (a) 4. (a) 5. (a) 6. (a) 7. (b) 8. (c) 9. (a) 10. (a)

II. Answer the following in One or Two sentences.(T.B.Page. 61.)

1. **State law of mass action.**

"The total mass of substances taking part in a chemical reaction remains the same throughout the change."

2. **State law of multiple proportion.**

"When two elements A and B combine to form two or more compounds, then different weights of B which combine with a fixed weight of A bears a simple numerical ratio to one another".

3. **Define law of conservation of mass.**

When two elements combine separately with a definite mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.

4. **Give the limitations of law multiple proportion.**

The law is valid till an element is present in one particular isotopic form in all its compounds. When an element exists in the form of different isotopes in its compounds, the law does not hold good.

III. Answer in brief. (T.B. Page 61)

1.State and explain Gay lussac law with a simple illustration.

Law: When two or more gases react with one another, their volumes bear simple whole number ratio with one another and to the volume of products (if they are also gases) provided all volumes are measured under identical conditions of temperature and pressure.

The law can be understood with the help of following example.

(i) Gaseous hydrogen and gaseous chlorine react together to form gaseous hydrogen chloride according to the following equation.

$$\underset{\text{One volume}}{H_2(g)} + \underset{\text{One volume}}{Cl_2(g)} \longrightarrow \underset{\text{Two volume}}{2\,HCl(g)}$$

It has been observed experimentally that in this reaction, one volume of hydrogen always reacts with one volume of chlorine to form two volumes of gaseous hydrogen chloride. All reactants and products are in gaseous state and their volumes bear a ratio of 1: 1: 2. **This ratio is a simple whole number ratio**.

(ii) Similarly, under suitable conditions, gaseous nitrogen and gaseous hydrogen combine together to form gaseous ammonia according to the equation

$$N_2(g) + 3H_2(g) \longrightarrow 2\,NH_3(g)$$

It has been found that one volume of nitrogen always reacts with three volumes of hydrogen to form two volumes of gaseous ammonia. Thus, the volumes of reactants and products bear the ratio 1: 3: 2 **which is a simple whole number ratio.**

2.What is the present day position of law of conservation of mass?

Present Day Position of the Law of conservation of mass: This law is particularly not applicable to nuclear reactions where tremendous amount of energy is liberated. However for chemical reactions, the law of conservation of mass is adequate, since energy changes are comparatively small (*i.e.,* the change in mass is immeasurably small or negligible).

3.State and explain the law of constant proportion with an illustration.

Law: A pure chemical compound always contains the same elements combined together in the same definite (fixed or constant) proportions by weight, irrespective of its source or method of preparation. Therefore, the law is also called as the law of fixed proportions or constant proportions.

ILLUSTRATIONS:

(a) Carbon dioxide may be obtained by the following methods:

(i) by burning carbon

(ii) By reaction between a metal carbonate and a dilute acid.

(iii) By heating calcium carbonate or sodium bicarbonate. Analysis of carbon dioxide, prepared by any of the above methods, shows that it contains only carbon and oxygen, combined together in the same proportion by weight, *i.e.,* 12: 32 or 3: 8.

IV. Answer in detail.

1.State and experimentally verify the law of conservation of mass

Law: Whenever a chemical change occurs, the total mass of products is the same as the total mass of reactants. Alternatively the law can be stated as **"the total mass of substances taking part in a chemical reaction remains the same throughout the change."**

Experimental verification of the law of conservation of mass: This law can be verified by the study of any chemical reaction. In the laboratory, it can easily be verified by the study of the following reaction.

$$AgNO_3 + NaCl \longrightarrow AgCl\downarrow + NaNO_3$$

When a solution of silver nitrate ($AgNO_3$) is treated with a solution of sodium chloride, a white precipitate of silver chloride (AgCl) is obtained along with a solution of sodium nitrate ($NaNO_3$). If the law is true, the total mass of $AgNO_3$ and NaCl should be the same as the total mass of AgCl precipitate and $NaNO_3$ solution. The experiment is done in a specially designed H shaped tube called Landolt's tube.

Sodium chloride solution is taken in one limb of the tube while silver nitrate solution is taken in the other limb as shown in the figure 3.1. Both the limbs are now sealed and tube is weighed. Now the tube is inverted so that the solutions can mix up together and react chemically. The reaction takes place as mentioned above and a precipitate of silver chloride is obtained. The tube is again weighed. The mass of the tube is found to be exactly the same as the mass obtained before inverting the tube. This experiment clearly shows that the law of conservation of mass is true.

Fig. Landolt's tube

Present Day Position of the Law of conservation of mass: This law is particularly not applicable to nuclear reactions where tremendous amount of energy is liberated. However for chemical reactions, the law of conservation of mass is adequate, since energy changes are comparatively small (*i.e.,* the change in mass is immeasurably small or negligible).

2.Give the experimental verification of law of constant composition.

Prepare pure samples of cupric oxide by two different methods.

(i) by heating copper carbonate

(ii) by the decomposition of cupric nitrate.

The cupric oxide prepared by both the methods always contains the same elements copper and oxygen combined together in the same fixed proportion of 4: 1 by weight. This illustrates the law of definite proportions. It can be verified by taking a known weight of a pure sample (W_1 gm) of cupric oxide in a porcelain boat.

$$2\,Cu(NO_3)_2 \longrightarrow 2\,CuO + 4\,NO_2\uparrow + O_2\uparrow$$

It is placed inside a hard glass tube kept horizontally as shown in fig 3.2.

A current of pure dry hydrogen is sent inside the tube and the tube is heated. The cupric oxide is reduced to metallic copper.

$$CuO + H_2 \longrightarrow Cu + H_2O$$

The Weight of copper formed is found out W_2gm.

Calculation:

Method 1:

$$\begin{aligned} \text{Weight of cupric oxide} &= W_1\text{gm} \\ \text{Weight of Copper} &= W_2\text{gm} \\ \therefore\ \text{Weight of oxygen} &= W_1 - W_2\ \text{gm} \\ \text{Ratio of copper : oxygen} &= W_2 : (W_1 - W_2) \end{aligned}$$

Fig. Verification of the Law Definite Proportion

The **same experiment is repeated** with a known weight W_3 gm of cupric oxide prepared by heating copper carbonate

$$CuCO_3 \longrightarrow Cu\,O + CO_2 \uparrow$$

The cupric oxide formed is reduced to metallic copper by passing a current of pure and dry hydrogen inside the tube as before. The weight of metallic copper was found to be W_4 gm. The ratio of the weight of copper to the weight of oxygen in both the samples are calculated as follows:

Method 2:

Weight of cupric oxide	=	W_3gm.
Weight of copper	=	W_4gm
$\therefore$ Weight of oxygen	=	$W_3 - W_4$ gm
Ratio of copper to oxygen	=	$W_4 : (W_3 - W_4)$

The two ratios $[W_2 : W_1 - W_2]$ and $[W_4 : W_3 - W_4]$ are found to be the same and is equal to 4: 1. ***Thus the law of definite proportions is verified experimentally***.

3.Explain the law of multiple proportions with suitable illustrations.

Law: "When two elements A and B combine to form two or more compounds, then different weights of B which combine with a fixed weight of A bears a simple numerical ratio to one another".

Explanation: Carbon combines with oxygen to form two different oxides, namely, carbon monoxide (CO) and carbon dioxide (CO_2). The proportions by weight of the two elements are

Carbon monoxide	-	C: O:: 12 : 16
Carbon dioxide	-	C: O:: 12 : 32

There, the weights of oxygen that combine with a fixed weight of carbon (12g) are in the ratio 16g : 32g *i.e.* 1:2, a simple numerical ratio.

Illustrations:

Nitrogen combines with oxygen to form different oxides. The compositions by weight of these oxides are shown in table.

Compositions by weight of oxides of nitrogen

No.	Name of Oxide	Wt. of nitrogen in grams	Wt. of oxygen in grams
1	Nitrous oxide (N_2O)	28	16
2	Nitric oxide (2NO)	28	32
3	Nitrogen trioxide (N_2O_3)	28	48
4	Nitrogen tetraoxide (N_2O_4)	28	64
5	Nitrogen pentoxide (N_2O_5)	28	80

It can be seen from the table that different weights of oxygen that combines with a fixed weight of nitrogen (28 g) are in the ratio,16g : 32 g : 48g : 64g : 80g *i.e.,* in the simple numerical ratio of 1 : 2 : 3 : 4 : 5.

4.Give the experimental verification of law of multiple proportions.

The law can easily be verified by the study of oxides of copper. Copper reacts with oxygen to form two oxides - the red cuprous oxide (Cu_2O) and the black cupric oxide (CuO).

In order to verify the law of multiple proportion, fixed amounts of these oxides (say 20g each) are separately reduced to metallic copper by heating them in a current of hydrogen and the masses of copper obtained from them are estimated. The difference in the mass of oxide taken and the mass of copper obtained from it gives the mass of oxygen present in it.

Now the masses of oxygen which combine with a definite mass of copper in the two oxides are calculated. These masses are found in a simple whole number ratio. This verifies the law of multiple proportion.

Present day position of Law of Multiple Proportion: The law is valid till an element is present in one particular isotopic form in all its compounds. When an element exists in the form of different isotopes in its compounds, the law does not hold good.

5.State and explain the law of reciprocal proportions.

Law: When two elements combine separately with a definite mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.

Illustrations:

1. Let us consider three elements hydrogen, sulphur and oxygen. Hydrogen combines with oxygen to form H_2O whereas sulphur combines with it to form SO_2. Hydrogen and sulphur can also combine together to form H_2S.

   The formation of these compounds is shown in figure.

Fig. Illustration of Law of reciprocal proportions

In H_2O, the ratio of masses of H and O is 2: 16.

In SO_2 , the ratio of masses of S and O is 32: 32.

Therefore, the ratio of masses of H and S which combine with a ***fixed mass of oxygen*** (say 32 parts) will be

4: 32 i.e., 1 : 8 .... *(1)*

When H and S combine together, they form H_2S in which the ratio of masses of

H and S is

2:32 *i.e.,* 1: 16 **.........** *(2)*

The two ratios *(i)* and *(ii)* are related to each other as

: or 2 : 1

i.e., they are whole number multiple of each other.

Thus, the ratio of masses of H and S which combines with a fixed mass of oxygen is a whole number multiple of the ratio in which H and S combine together.

Sulphur and oxygen combine together to form SO_3 also. This case can also be worked out in the same way as above and can be shown to follow the law of reciprocal proportions.

V. Problems.(T.B. Page 61 –62)

1.In an experiment 5.0g of $CaCO_3$ on heating gave 2.8 g of CaO and 2.2 g of CO_2. Show that these results are in accordance to the law of conservation of mass.

Solution: $CaCO_3$ $CaO + CO_2$

Weight of $CaCO_3$ = 5.0 gms

Weight of CaO = 2.8 gms

Weight of CO_2 = 2.2 gms

Total weight of reactant = Total weight of products.

5.0 = 2.8 + 2.2

5.0 = 5.0

Since, the mass of the reactants are equal to the mass of the product, these results are in accordance to the laws of conservation of mass.

2.In an experiment 48 gms of magnesium combines with 32 gms of oxygen to form 80 gms of magnesium oxide. Show that this reaction illustrates the Law of Conservation of Mass.

[Hint: $2Mg + O_2 \longrightarrow 2MgO$. Atomic mass of Mg = 24 and O = 16].

Solution: $2\,Mg + O_2 \rightarrow 2MgO$

Magnesium + oxygen → Magnesium oxide

Weight of Magnesium = 48 gms

Weight of oxygen = 32 gms

Weight of magnesium oxide = 80 gms

∴ Total weight of reactants = Total weight of products

48 + 32 = 80

80 gms = 80 gm.

So these results are in accordance to the laws of conservation of mass.

3. 1.375 g of CuO were reduced by H_2 and 1.098 g of Cu were obtained. In another experiment, 1.178 g of Cu were dissolved in nitric acid and the resulting copper nitrate was converted into CuO by ignition. The weight of CuO formed was 1.476 g. Show that these results prove the law of constant proportion.

Solution:

Experiment 1:

Weight of CuO = 1.375 g
Weight of Cu = 1.098 g
Weight of oxygen = (1.375 – 1.098) g
= 0.277 g
Ratio of copper oxygen = 1.098 : 0.277
= 3.96 : 1

Experiment 2:

Weight of CuO = 1.476
Weight of Cu = 1.178
Weight of oxygen = 1.476 – 1.178
= 0.298
Ratio of copper: oxygen = 1.178 : 0.298
= 3.96 : 1

In both experiments the ratio of Copper: oxygen is some (3.96: 1). Hence it illustrates the law of definite proportions.

4.In an experiment 0.2430 gm of magnesium on burning with oxygen yielded 0.4030 gm of magnesium oxide. In another experiment 0.1820 gm of magnesium on burning with oxygen yielded 0.3020 gm of magnesium oxide. Show that the data explain the law of definite proportions.

Solution:

Experiment 1:

Weight of Magnesium oxide = 0.4030 gm
Weight of Magnesium = 0.2430 gm
Weight of oxygen = 0.4030 – 0.2430
= 0.16 gm
Ratio of Magnesium: oxygen = 0.2430: 0.16
= 1.552: 1

Experiment 2:

Weight of Magnesium oxide = 0.3020
Weight of Magnesium = 0.1820
Weight of oxygen = 0.3020 – 0.1820
= 0.12
Ratio of magnesium: oxygen= 0.1820: 0.12
= 1.552: 1

In both experiments the ratio of magnesium: oxygen is same (1.518:1) Hence it illustrates the law of definite proportions.

5. In an experiment 34.5 g oxide of a metal was heated so that O_2 was liberated and 32.1 g of metal was obtained. In another experiment 119.5 g of another oxide of the same metal was heated and 103.9 g metal was obtained and O_2 was liberated. Calculate the mass of O_2 liberated in each experiment. Show that the data explain the law of multiple proportions.

Solution: Experiment 1

Weight of the metal oxide = 34.5 g
Weight of the metal = 32.1 g

32.1 g metal combines with 2.4 g oxygen.

1 g of the metal combine with $\frac{2.4}{32.1}$ = 0.075 g

Experiment 2

Weight of the oxide taken = 119.5 g

Weight of the metal formed = 103.9 g

Weight of oxygen liberated = 15.6 g

103.9 g of metal combines with 55.6 g oxygen.

$$1 \text{ g of metal} = \frac{15.6}{103.9} \times 1 = 0.15014 \text{ oxygen}$$

Therefore different weights of oxygen, that combine with the fixed weight of the metal viz 1 g are in the ratio

0.1501 : 0.075

2 : 1

6.Copper combines with oxygen to form two oxides, which have the following composition:

***(i)* 0.716 g of cuprous oxides contains 0.630 g of copper.**

***(ii)* 0.398 g of cupric oxide contains 3.318 g of copper.**

Prove that the above data illustrates the law of multiple proportions.

Solution: Here Copper forms different oxides

	Cuprous oxide(A)	**Cupric oxide (B)**
Weight of oxide Weight of copper Weight of oxygen	0.716 g 0.630 g 0.086 g	0.398 g 0.318 g 0.08 g

Thus 0.630, 0.318 gram of copper combines with definite weight of oxygen

0.080: 0.080 = 1 : 1

The proportion by weight of chlorine is indicated by simple ratio. Thus law of multiple proportions is obeyed.

7. One gram of hydrogen combines with 15.88 g of sulphur. One gram of hydrogen combines with 7.92 g of oxygen, one gram of sulphur combines with 0.998 g of oxygen. Show that these data illustrate the law of reciprocal Proportions.

Solution: In hydrogen – Sulphur and Hydrogen – Oxygen combinations

The weight of Hydrogen = 1.0 grams

Weight of sulphur = 15.88 grams

Weight of oxygen = 7.92 grams

In Hydrogen – oxygen combinations the ratio masses of H and O is 1:8 ..... *(1)*

In Hydrogen – sulphur combinations the ratio masses of H and S is 1:16 ..... *(2)*

So the ratio *(1)* and *(2)* are related to each other as

$$\frac{1}{8} : \frac{1}{6} \text{ (or)} \qquad \ldots\ldots (3)$$

They are whole no multiple of each other.

In oxygen – sulphur combinations.

Weight of sulphur = 1 gram

Weight of oxygen = 0.998 gram

The ratio masses of H and S is 1:1 ..... *(4)*

(1), *(2)* and *(3)* are simple multiples of each other therefore, the law of reciprocal proportions holds good.

8. A compound contains carbon and chlorine. The percentage of chlorine in the compound is 92.21. In another compound which contains carbon and sulphur, the percentage of sulphur is 84.21. In a third compound which contains sulphur and chlorine, the percentage of chlorine is 52.59. Show that these data illustrate the law of reciprocal proportions.

Solution: In first compound (is carbon and chlorine)

The weight of chlorine = 92.21 gram

The weight of oxygen = (100 – 92.21)

= 17.79 gram

9.Calculate the volume of oxygen required for the complete combustion of $20cm^3$ of methane. ***[Hint: $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$]***

Solution:

CH_4	+	$2O_2 \longrightarrow$	CO_2	+	$2H_2O$
1 volume		2 volumes	1 volume		2 volumes
$20\ cm^3$		$2 \times 20cm^3$	$20\ cm^3$		$40\ cm^3$

Volume of O_2 required for combustion = $40cm^3$

10. $100cm^3$ of propane was burnt in excess oxygen to form carbon dioxide and water. Calculate: *(i)* the volume of oxygen used up, and *(ii)* the volume of carbon dioxide formed. ***[Hint: $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O$].***

Solution:

$C_3H_{8(g)}$	+	$5O_{2(g)} \longrightarrow$	$3CO_{2(g)}$	+	$4H_2O$
1 volume		5 volumes	3 volume		4 volumes
$100\ cm^3$		$500\ cm^3$	$300\ cm^3$		$400\ cm^3$

Volume of O_2 used = $500cm^3$

Volume of CO_2 formed = $300\ cm^3$

TEXT BOOK PROBLEMS (PAGES 52, 55, 56, 58& 59)

Problem 1. 1.5g mercury when heated in air, produced 1.62 g mercury oxide, 2.16g mercury oxide, on the other hand on heating in air produced 0.16g oxygen, leaving behind 2.0g mercury. Show that these results are according to the law of definite proportions.

Solution:

(i) 1.62g oxide – 1.5g mercury = 0.12 g oxygen.

1.5g mercury combines with = 0.12g oxygen.

$\therefore$ 1.0g mercury will combine with = $\dfrac{0.12 \times 1}{1.5}$ = 0.08 g oxygen.

(ii) 2.16g oxide –2.0 g mercury = 0.16 g oxygen.

2.0 g mercury combines with = 0.16g oxygen.

$\therefore$ 1.0g mercury will combine with = $\dfrac{0.16 \times 1}{2}$ = 0.08 g oxygen.

In first case oxide was produced, in second case oxide was decomposed. Both show that 1.0g mercury has combined with 0.08 g oxygen. That means, proportion of mercury to oxygen is 1: 0.08, in both the cases. **This supports the law of definite proportions**.

Problem 2: 1.19g of Zinc was converted into zinc oxide and 1.51 gm of zinc oxide was obtained. In another experiment 1.812 gm of zinc oxide when heated with carbon gave 1.428 gm of zinc show that these results illustrate the law of definite proportions.

Solution:

Experiment 1:

Wt. of Zinc oxide = 1.51 gm

Wt. of Zinc = 1.19 gm

$\therefore$ Wt of oxygen = 1.51 - 1.19

= 0.32 gm.

Ratio of Zinc: oxygen = 1.19: 0.32

= **3.72: 1**

Experiment 2:

Wt of zinc oxide = 1.812 gm

Wt of Zinc = 1.428 gm

$\therefore$ Wt of oxygen = 1.812 - 1.428

= 0.384 gm

Ratio of Zinc: oxygen = 1.428: 0.384

= **3.72: 1**

In both the experiments the ratio of Zinc: oxygen is same (3.72:1). Hence it illustrates the law of definite proportions.

Example: 1.(P.55) Iron forms two different chlorides, namely ferrous and ferric chloride. Each of these chlorides was prepared from 2.0 g iron. It was found that 4.538 g ferrous and 5.804g ferric chloride were produced. Show that these observations are according to the law of multiple proportions.

Solution: Here iron is forming different chlorides. The weight of iron taken in both cases is the same i.e. 2.0g. Therefore we have,

Ferrous chloride (A)		Ferric chloride (B)	
4.538 g	chloride	5.804 g	chloride
-2.000 g	iron	- 2.000 g	iron
----------		------------	
2.538 g	chlorine	3.804 g	chlorine.

Thus a definite weight of iron i.e.2.0g, combines with 2.538 g and 3.804 g chlorine. The proportion of chlorine in these compounds is

Ferrous : Ferric

2.538 : 3.804 = 1:1.5 = 2:3

The proportion by weight of chlorine is indicated by a simple ratio, ***Thus law of multiple proportions is obeyed.***

Example 2: Lead forms three oxides A, B and C. The quantity of oxygen in each of the oxides. A, B and C is 7.143 %, 10.345% and 13.133% respectively. Show that the law of multiple proportions is obeyed.

Solution: As the % of oxygen is given, we can find the % of lead.

	A	B	C
Oxide	100.000	100.000	100.000
-Oxygen	-7.143	-10.345	-13.133
-----------	-------------	-------------	-------------
Lead	92.857	89.655	86.867

As lead is forming different oxides, **let us take a fixed wt. of lead say 10 g, and find out the weights of oxygen combining with 10g lead in three oxides**.

In A, 92.857 g lead combines with = 7.143 g oxygen.

$\therefore$ 10.0 g lead will combine with $= \dfrac{7.143}{92.857} \times \dfrac{10}{1} = 0.769$ g oxygen.

Similarly, in B, 10 g lead will combine with $= \dfrac{10.345}{89.655} \times \dfrac{10}{1} = 1.054$ g oxygen

Finally in C, 10 g lead will combine with $= \dfrac{13.133}{86.867} \times \dfrac{10}{1} = 1.538$ g oxygen

The weights of oxygen combining with fixed wt. of lead (10.0g) are in the proportions,

0.769: 1.154: 1.538 = 1: 1.5: 2

i.e. 2: 3: 4

The proportion by weight of oxygen is given by a simple numerical ratio. ***Thus law of multiple proportions is obeyed.***

Problem 1(p 58) : Hydrogen sulphide (H_2S) contains 94.11% sulphur, water (H_2O) contains 11.11% hydrogen and sulphur dioxide (SO_2) contains 50% oxygen. Show that the results are in agreement with the law of reciprocal proportions.

Solution: In water, the weight of hydrogen = 11.11 g

The weight of oxygen = 100 -11.11 = 88.89 g

In sulphur dioxide, the weight of sulphur = 50g

The weight of oxygen = 100 -50 = 50g

$\therefore$The weight of sulphur that combines with 88.89 g of oxygen $= \dfrac{50}{50} \times 88.89 = 88.89$ g

∴ The ratio between the weights of sulphur and hydrogen which combine with a fixed weight of oxygen (88.89 g) is 88.89**:** 11.1 or 8:1. **...(i)**

In hydrogen sulphide, the weight of sulphur = 94.11g

The weight of hydrogen = 100 -94.11 = 5.89g

∴ The ratio between weights of sulphur and hydrogen is 94.11 **:** 5.89 or 16:1 **...(ii)**

The two ratios (i) and (ii) are related as $\frac{8}{1} : \frac{16}{1}$ **or 1 :2**

Which are simples multiples of each other. Therefore, the law of reciprocal proportions holds good.

Problem 1(p 59): Methane burns in oxygen to form carbon dioxide and water vapour as given by the equation

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Calculate: i) the volume of oxygen needed to burn completely 50 cm^3 of methane and (ii) the volume of carbon dioxide formed.

Solution:

CH_4	+	$2O_2$	$\longrightarrow$	CO_2	+	$2H_2O$
1 vol		2vols		1vol.		2vols.
1 x 50 cm^3		2x50 cm^3		1 x50cm^3		2 x 50 cm^3
50cm^3		100cm^3		50cm^3		100cm^3

Volume of oxygen used = 50 cm^3

Volume of carbon dioxide formed = 50 cm^3

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1. The following is the best example of law of conservation of mass:
   (a) 12g of carbon is heated in vacuum, there is no change in mass.
   (b) The weight of wire of platinum is the same before and after heating.
   (c) A sample of air increases in volume when heated at constant pressure but remains unchanged.
   (d) 12g of carbon combines with 32g of oxygen to give 44g of carbon dioxide.
2. The formation of CO and CO_2 illustrates the law of:
   (a) Conservation of mass (b) Constant proportion
   (c) Multiple proportion (d) Reciprocal proportion
3. An element X forms two oxides containing 53.33% and 36.36% of oxygen, respectively. These data illustrate the law of:
   (a) Constant Proportion (b) Conservation of mass
   (c) Multiple proportion (d) Reciprocal proportion
4. Hydrogen sulphide contains 5.88% hydrogen, H_2O contains 11.11% hydrogen while SO2 contains 50% suphur. These figures illustrate the law of:
   (a) Conservation of mass (b) Constant proportion
   (c) Multiple proportion (d) Reciprocal proportion
5. 2g of H_2 combine with 16g of O_2 to form H_2O and 60g of carbon to form CH_4. In CO_2, 12g of C are combined with 32g of O_2. These data illustrate the law of:
   (a) Reciprocal proportion (b) Multiple proportion
   (c) Constant Proportion (d) Conservation of mass
6. 143g $AgNO_3$ reacts with NaCl solution and gave 143.5g AgCl and 85g $NaNO_3$. If the law of conservation of mass holds, the weight of NaCl will be:
   (a) 5.85g (b) 58.5g (c) 117.0g (d) 11.70g
7. Phosphorus forms three oxides containing 39.22, 49.22 and 56.40 percent phosphorus. This example illustrates the law of:
   (a) Constant Proportion (b) Multiple proportion
   (d) Reciprocal proportion (d) Conservation of mass
8. 1g of chloride of iron gave 2.26 of AgCl while 1g of another chloride of iron 2.65 of AgCl. These result illustrate the law of:
   (a) Conservation of mass (b) Multiple proportion
   (c) Constant Proportion (d) Reciprocal proportion
9. Law of constant composition is same as the law of ?
   (a) Conservation of mass (b) conservation of energy
   (c) Multiple proportion (d) Definite proportion
10. Which of the following pairs of substances illustrates the law of multiple proportions?
    (a) CO and CO_2 (b) Nacl and NaBr (c) H_2O and D_2O (d) Mg_2O and Mg $(OH)_2$
11. If water samples are taken from sea, rivers, clouds, lakes or show, they will be found to contain hydrogen and oxygen in the approximate ratio of 1:8. This indicates the law of?
    (a) Multiple proportions (b) Definite proportions (c) Reciprocal proportions (d) None

12. Law of multiple proportions is illustrated by the following pair of compounds?
(a) H_2S and SO_2 (b) N_2O and NO (c) HCl and HNO_3 (d) KOH and KCl

13. In SO_2 and SO_3 the ratio of the weights of oxygen which combines with a fixed weight of sulphur is
13. This is an example of law of?
(a) Constant proportion (b) Multiple proportion (c) Reciprocal proportion (d) Gay-Lussac

14. The law of conservation of mass holds good for all of the following except,
(a) All chemical reactions (b) Nuclear reactions
(c) Endothermic reactions (d) Exothermic reactions

15. The oxide of nitrogen contain 63.65%, 46.69% and 30.45% nitrogen respectively. The data illustrate the law of?
(a) Definite proportion (b) Multiple proportion
(c) Reciprocal proportion (d) Conservation of mass

16. Water and hydrogen peroxide illustrate the law of?
(a) Reciprocal proportions (b) Multiple proportions
(c) Constant proportions (d) None

17. The percentage of hydrogen in water and hydrogen peroxide are respectively 11.2% and 5.94%. This illustrates?
(a) Law of multiple proportions (b) Conservation of mass
(c) Law of definite proportions (d) Law of reciprocal proportions

18. Which of the following illustrates the law of multiple proportions?
(a) H_2O, Na_2P (b) MgO, Na_2O (c) Na_2O, BaO (d) $SnCl_2$, $SnCl_4$

19. A chemical equation is balanced in accordance with the law of?
(a) Constant proportion (b) Multiple proportion
(c) Reciprocal proportion (d) Conservation of mass

20. H_2O contains 88.8% oxygen and 11.2% Hydrogen by weight which illustrate the law of
(a) Multiple proportions (b) Definite proportion
(d) Combining volumes (d) Reciprocal proportion

21. Which of the following best explains the law of conservation of mass?
(a) No change in mass is observed when 2.0g of Mg is heated in vacuum,
(b) 1.2g of carbon when burnt in excess of oxygen consumes only 3.2g of it to form 4.4 g of carbon dioxide.
(c) 12g of carbon when heated in limited supply of air produces only 20g of carbon monoxide
(d) A sample of air on heating does not show any change in mass but volume increases.

22. Which of the following set illustrated law of reciprocal proportions?
(a) PCl_3, HCl, HBr (b) PH_3, P_2O_3, P_2O_5 (c) PH_3, P_2O_3, H_2S (d) PH_3, P_2O_5, H_2O.

23. Which of the following illustrates the law of conservation of mass?
(a) Mixing of 10g of sulphur and 2g of sand does not show a change in mass
(b) The mass of platinum wire before and after heating remains constant
(c) 2.2g of propane and 8g of oxygen produces 10.2g of gaseous mixture
(d) 2.8g of CO and 1.6g of oxygen gave only 2.24L of CO2 at S.T.P.

24. Which of the following is the best example of law of conservation of mass:
(a) 12g of carbon combines with 32g of oxygen to form 44g of carbon dioxide.
(b) 12g of carbon is heated in vacuum, there is no change in mass.
(c) The mass of a piece of platinum is the same before and after heating.
(d) A sample of air increases in volume when heated at constant pressure but the mass remains unchanged.

25. In compound A, 1-0g N2 unites with 0.57g O2. In compound B,2g N2 combines with 2.25g O2. In compound C, 3.0g N2 combines with 5.11g O2. These results obey the law of:
(a) Law of constant of mass (b) Law of multiple proportions
(c) Law of reciprocal proportions (d) Dalton's law of partial pressure.

26. The formation of CO and CO2 illustrate the law of:
(a) Conservation of mass (b) Constant proportion
(c) Multiple proportion (d) Reciprocal proportion

27. In the following reaction:
$H_2 + Cl_2 \rightarrow$ 2HCl the ratio of volumes of H2, Cl_2 and HCl gas is 1:1:2. These figures illustrate the law of:
(a) Constant proportion (b) Multiple proportion
(c) Reciprocal proportion (d) Gay Lussac's law of gaseous volume.

28. H_2S contains 5.88% hydrogen, H2O contains 11.11% hydrogen while SO_2 contains 5o% s. These figures illustrate the:
(a) Law of conservation of mass (b) Law of constant proportion
(c) Law of multiple proportion (d) Law of reciprocal proportion

29. An element X forms two oxides containing 53.33% and 36.36 % of oxygen respectively. These data illustrate the law of:
(a) Conservation of mass (b) Constant proportions
(c) Multiple proportions (d) Reciprocal proportions

30. CO_2 gas was prepared by (i) strongly heating $NaHCO_3$. (ii) burning charcoal in air and (iii) the action of $CaCO_3$ and dil. HCl. It was found that in each case carbon and oxygen combined in the ratio of 3:8. These data illustrate the law of:
(a) Conservation of mass (b) Constant proportions
(c) Multiple proportions (d) Reciprocal proportions
31. 1.0 gram of an oxide of A contained 0.5 gram of A, 4.o gram of another oxide of A contained 1.6g. Of A. These figures illustrate the:
(a) Law of reciprocal proportion (b) Law of multiple proportion
(c) Law of conservation of mass (d) Law of conservation proportion.
32. (i) In hydrogen chloride H = 2.77% and Cl = 97.23% (ii) In phhosphine P =91.18% and H = 8.82% (iii) In phosphorous pentachloride P= 22.57% and Cl = 77.43%
(a) Reciprocal proportion (b) Multiple proportion
(c) Constant Proportion (d) Conservation of mass
33. Nitrogen forms five stable oxides with oxgen of formulae, NO, N_2O, N_2O_3, N_2O_4, N_2O_5. The formation of these oxides explain fully the
(a) law of definite proportions (b) law of partial pressure (c) law of multiple proportion
34. The percentage of CO and S in CO_2, SO_2 and CS_2 when considered together illustrates the law of
(a) Constant proportion (b) Multiple proportion
(c) Reciprocal proportion (d) Conservation of mass.
35. In the reaction, $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio of volumes of nitrogen hydrogen and ammonia is 1:3:2: These figures illustrate the law of
(a) Constant proportion (b) Multiple proportion
(c) Reciprocal proportion (d) Gay Lussac's law of gaseous volume.
36. The percentage of sliver and chlorine in two samples of silver chloride prepared b heating silver foil in the current of chlorine and by the interaction of silver nitrate and hydrochloric acid were found to be identical. This illustrates the law of
(a) Conservation of mass (b) Constant proportions
(c) Multiple proportions (d) Reciprocal proportions
37. Which of the following statements is not corrected?
(a) One mole of carbon and 1/3 mole of carbon dioxide contain same number of atoms.
(b) One mole of NH_3 and one mole of BF_3 contains same number of atoms.
(c) One mole of CO_2 occupies more volumes than one mole of CO at N.T.P.
(d) One mole of sodium hydroxide base 20 g
38. Law of definite proportion or constant composition was stated by …………….
(a) J.L. Proust (b) Lavoisier (c) Wenzel (d) Gay Lussac
39. Law of multiple proportion was stated in the year …………..
(a) 1756 (b) 1774 (c) 1799 (d) 1803
40. Law of multiple proportion was experimentally verified and confirmed by Berzelius in the year ……..
(a) 1803 (b) 1800 (c) 1811 (d) 1799
41. The weights of oxygen that combine with a fixed weight of carbon in CO and CO_2 are in the ratio ……….
(a) 16: 32 or 1:2 (b) 1:3 (c) 1: 4 (d) 4: 1
42. The ratio of nitrogen and oxygen by weight in N_2O and NO is …………
(a) 1:2 (b) 2:1 (c) 2:3 (d) 1:4
43. The simple numerical ratio between nitrogen and oxygen in its various oxides namely N_2O, 2NO, N_2O_3, N_2O_5 is ………….
(a) 2:3:4:5:1 (b) 1:3:4:2:5 (c) 1:2:3:4:5 (d) 4:3:2:1:5
44. Law of reciprocal proportion was stated in the year …………..
(a) 1812 (b) 1777 (c) 1792 (d) 1803
45. 1 cm^3 = …………….
(a) 10 ml (b) 0.1 ml (c) 1 ml (d) 0.01 ml
46. The ratio of the volume of reactants and products in the formation of HCl is …………..
(a) 1:2:1 (b) 1:1:2 (c) 1:1:1 (d) 2:1:1
47. Law of combining volumes was stated by Gay Lussac in the year ………..
(a) 1803 (b) 1805 (c) 1800 (d) 1808
48. The ratio of masses of S and O in SO_2 molecule is ………….
(a) 32:16 (b) 16:32 (c) 32:32 (d) None
49. In H_2O, the ratio of masses of H and O is …………
(a) 2:16 (b) 1:16 (c) 16:2 (d) 16:1
50. Richter carried out some basic experiments on the law of reciprocal in the year ………………
(a) 1777 (b) 1792 (c) 1803 (d) 1775
51. Which law does not holds good for the element exists in different isotopic forms?
(a) Law of multiple proportions (b) Law of conservation of mass
(c) Law of definite proportions (d) Law of reciprocal proportions
52. Chlorine combines with hydrogen and carbon to form HCl and CCl_4 when hydrogen and carbon combines they do so in a ratio ………..
(a) 1:3 (b) 1:12 (c) 36.45:12 (d) 4:12
53. Sodium combines with isotopes of ${}_{17}Cl^{35}$ and ${}_{17}Cl^{35}$ to form sodium chloride. Their formation illustrates…
(a) the law of definite proportions (b) the law of multiple proportions

(c) the law of reciprocal proportions (d) None of these

54. Ammonia is formed by the reaction, $N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$. The volume of ammonia formed from 6 litres of nitrogen is ……………..
(a) 12 litres (b) 18 litres (c) 6 litres (d) 2 litres

55. 36.45 g of HCl reacts with 40 g of sodium hydroxide to form sodium chloride and 18 g of water. The mass of sodium chloride is ……………….
(a) 58.46 g (b) 40 g (c) 117 g (d) 36.45 g

56. 2 g of H_2 combines with 16g of O_2 to form H_2O and with 6g of C to form CH_4. In CO_2, 12 g of C are combined with 12 g of O_2. These data illustrate the law of ……………….
(a) reciprocal proportion (b) multiple proportion (c) constant proportion (d) conservation of mass

57. CO_2 was prepared by (i) strongly heating $NaHCO_3$, (ii) burning charcoal in air and (iii) action of $CaCO_3$ and dil. HCl. It was found that in each case, carbon and oxygen combined in the ratio of 3:8. These data illustrate the law of ………………..
(a) conservation of mass (b) constant proportion (c) multiple proportion (d) reciprocal proportion

58. 1.0g of an oxide of A contained 0.5g of A, 4.0g of another oxide of A, contained 1.6g of A. These figures illustrate the ………………..
(a) law of reciprocal proportion (b) law of multiple proportion
(c) law of conservation of mass (d) law of constant proportion

59. The percentage of H in H_2O_2 and H_2O is 5.93 and 11.2 respectively. These figures illustrate the law of …………….
(a) Conservation of mass (b) Constant proportion (c) Multiple proportion (d) Reciprocal proportion

60. In hydrogen chloride H = 2.77% Cl = 97.23% In phosphine P = 91.18% and H = 8.82%
In phosphorous pentachloride P = 22.57% Cl = 77.43% these figures illustrate the law of ………..
(a) reciprocal proportion (b) multiple proportion (c) constant proportion (d) conservation of mass

61. A sample of $CaCO_3$ has Ca = 40%, C = 12% and O = 48%. If the law of constant proportions is true then the mass of Ca in 4g of a sample of $CaCO_3$ from another source will be …………..
(a) 0.016 g (b) 0.16g (c) 1.6 g (d) 16 g

62. The composition of compound A is 40% X and 60% Y. The composition of compound B is 25%. X and 75% Y. According to the law of multiple proportion the ratio of the mass of element Y in elements A and B is ……………….
(a) 4:5 (b) 2:1 (c) 2:3 (d) 3:4

63. Zinc sulphate contains 22.65% Zn and 43.9% H_2O. If the law of constant proportion is true, then the mass of Zn required to give 20g of the crystals will be …………..
(a) 0.453 g (b) 4.53 g (c) 45.3 g (d) 453 g

64. In the following reaction: $H_2 + Cl_2 \longrightarrow 2HCl$ the ratio of volumes of H_2, Cl_2, HCl gas is 1:1:2 these figures illustrate the law of ………………..
(a) constant proportion (b) multiple proportion
(c) reciprocal proportion (d) Gay Lusac's law of gaseous volume

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Answers:

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

14. (b) 15. (b) 16. (b) 17. (a) 18. (d) 19. (d) 20. (b) 21. (b) 22. (d) 23. (c) 24. (a) 25. (b) 26. (c)

27. (d) 28. (d) 29. (c) 30. (b) 31. (b) 32. (a) 33. (c) 34. (a) 35. (d) 36. (c) 37. (d) 38. (a) 39. (a)

40. (c) 41. (a) 42. (a) 43. (c) 44. (a) 45. (c) 46. (b) 47. (d) 48. (c) 49. **(a) 50. (b) 51. (a) 52. (b) 53. (d) 54. (a)**

55. (a) 56. (a) 57. (b) 58.(b) 59. (c) 60. (a) 61. (c) 62. (a) 63. (a) 64. (d)

II. Answer the following in one or two sentences:

1.List the various laws of chemical combinations.

1. Law of conservation of mass – stated by Lomonossoff
2. Law of definite proportions – stated by J.C. Proust
3. Law of multiple proportions – stated by Dalton
4. Law of Reciprocal proportions – stated by Beizelius
5. Gay Lussac's Law of combining volumes – stated by Gay Lussac's

2.What is stoichiometry?

Stoichiometry a branch of chemistry in which quantitative relationship between masses of reactants and products are established. The study of these laws led to the development of a theory concerning the nature of matter.

3. State law of reciprocal proportions.

Law: When two elements combine separately with a definite mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.

4. State Gay Lussac's law of combining volumes.

When two or more gases react with one another, their volumes bear simple whole number ratio with one another and to the volume of products (if they are also gases) provided all volumes are measured under identical conditions of temperature and pressure.

5. What is the present day position of the law of conservation of mass?

The law of conservation of mass is particularly not applicable to nuclear reactions where tremendous amount of energy is liberated. However for chemical reactions, the law of conservation of mass is adequate, since energy changes are comparatively small

6. State the law of definite proportions?

A pure chemical compound always contains the same elements combined together in the same definite proportions by weight, irrespective of its source or method of preparation. Therefore this law is also called the law of fixed (or) constant proportions.

ADDITIONAL PROBLEMS BASED ON LAWS OF CHEMICAL COMBINATIONS

I. *LAW OF CONSERVATION OF MASS*:

1. If 3.0 g of magnesium combine with 2.0 g of oxygen to form magnesium oxide, what weight of magnesium oxide could be formed from 4.8 g of magnesium? Show how these data illustrate law of conservation of mass.

Case 1

$$Mg + O_2 \rightleftarrows 2\,MgO$$

Weight of magnesium	=	3g
Weight of oxygen	=	2g
Weight of magnesium from which Magnesium oxide is to be formed	=	4.8g
Weight of magnesium oxide	=	?

According to **the law of conservation of mass, "**the mass of magnesium oxide formed should be equal to the total mass of magnesium and oxygen".

3 g of Mg gives	=	5 g of MgO
$\therefore$ 4.8 g of Mg will give	=	$\frac{5}{3}$ x 4.8 of MgO
	=	8 g of MgO
Weight of magnesium oxide	=	**8g**

2. 48 gms of magnesium combines with 32 gms of oxygen to form 80 gms of magnesium oxide. Show that his reaction illustrates the Law of Conservation of Mass.

(Mg = 24; O = 16)

Solution

$2Mg + O_2 \longrightarrow 2MgO$

Total mass of reactants ($2\,Mg + O_2$)	=	48 gms + 32 gms
	=	80 gms
Mass of product (2MgO)	=	80 gms
Total mass of reactants	=	Total mass of products.

Thus the reaction illustrates the Law of Conservation of Mass.

3. $2Cu + S \longrightarrow Cu_2S$. Show that this reaction illustrates the Law of conservation of Mass.

(Cu = 64, S = 32).

Solution

2Cu	+	S	$\longrightarrow$	Cu_2S
(2 x 64)	+	32	$\longrightarrow$	160
128	+	32	$\longrightarrow$	160

128 gms of copper + 32 gms of Sulphur $\longrightarrow$ 160 gms of cuprous sulphide.

i.e., 160 gms of reactants $\longrightarrow$ 160 gms of products.

i.e., Total mass of reactants = Total mass of products.
Thus the Law of Conservation of Mass is illustrated.

II. *LAW OF CONSTANT OR DEFINITE PROPORTION:*

1. In one experiment 1.098 g of copper is obtained by the reduction of 1.375 g of cupric oxide. In another experiment 1.476 g of cupric oxide was prepared from 1.179 g of copper through cupric nitrate. Show that the results of the two experiments illustrate the law of definite proportions?

Experiment: 1

Weight of curpric oxide = 1.375 g
Weight of copper = 1.098 g

Weight of oxygen = 0.277 g

Ratio in the weights of copper and oxygen = 1.098: 0.277 = **4:1**

Experiment: 2

Weight of the cupric oxide = 1.476 g
Weight of the copper = 1.179 g

Weight of oxygen = 0.297 g

Ratio in the weights of copper and oxygen = 1.179: 0.297
= **4:1**

In both cases, it is found that copper and oxygen are in the ratio **4:1,**
Hence, the law of definite proportion is proved.

2. Illustrate the law of definite proportions from the following data:
(i) 0.16 g of sulphur produces 0.32 g of sulphur dioxide, (ii) sulphur dioxide obtain by the decomposition of sodium sulphate contains 50% sulphur.

Case 1

(i) Weight of sulphur dioxide = 0.32 g
Weight of sulphur = 0.16 g

Weight of oxygen = 0.16 g

Ratio of sulphur to oxygen in sulphur dioxide produced by (i)
Method = **0.16: 0.16 or 1:1**

Case 2

(ii) Weight of sulphur dioxide = 100 g
Weight of sulphur = 50 g

Weight of oxygen = 50 g

Ratio of sulphur to oxygen in sulphur dioxide produced by (ii)
Method = **50:50 or 1:1**

Since the sulphur dioxide produced by different methods contains sulphur and oxygen in the ratio 1:1, the data proves **the law of definite proportions.**

3. In two experiments, 0.259 g and 0.207 g of lead were converted to lead chloride, yielding 0.347 g and 0.278 g of lead chloride respectively. Show that the data illustrates the law of constant composition.

Experiment: 1

Weight of lead chloride = 0.347 g
Weight of lead = 0.259 g

Weight of chlorine = 0.088 g

In this experiment, the lead and chlorine are present in the ratio **0.259: 0.088 or 3:1.**
In the second experiment

Experiment: 2

Weight of the lead chloride = 0.278 g

Weight of the lead	=	0.207 g

Weight of chlorine	=	0.071 g

The lead and chlorine are present in the ratio **0.207: 0.071 or 3:1.**

Since in both the experiments, it is found that in lead chloride, lead and chlorine are present in the ratio 3:1. This proves **the law of constant composition.**

4. In a typical experiment, 28 gms. of iron on heating with oxygen gave 40 gms. of iron (III) oxide. 7 gms. of iron on heating with oxygen gave 10 gms. of iron (III) oxide. Show that these results correspond to the law of definite proportions.

Solution: Case (i)

Weight of iron (III) oxide	=	40 gms.
Weight of iron	=	28 gms.
∴Weight of oxygen	=	12 gms.
Weight of iron: Weight of oxygen	=	28 : 12
	=	7 : 3

Case (ii)

Weight of iron (III) oxide	=	10 gms.
Weight of iron	=	7 gms.
∴Weight of oxygen	=	3 gms.
Weight of iron: Weight of oxygen	=	7 : 3

In both the cases, the ratio of weight of iron: weight of oxygen remains the same as 7:3. Thus these results correspond to the law of definite proportions.

5. In an experiment, 9 gms. of magnesium gave 15 gms. of magnesium oxide. 4.5 gms. of magnesium oxide was obtained from 2.7 gms. of magnesium. Show that these information's illustrate the law of definite proportions.

Solution: Case (i)

Weight of MgO	=	15 gms.
Weight of Mg	=	9 gms.
Weight of O	=	6 gms.
Weight of Mg: Weight of O	=	9 : 6 = 3: 2

Case (ii)

Weight of MgO	=	4.5 gms.
Weight of Mg	=	2.7 gms.
Weight of O	=	1.8 gms.
Weight of Mg: Weight of O	=	2.7: 1.8 = 3: 2

In both the cases, the ratio of weight of magnesium and weight of the oxygen mains the same as 3:2. Hence it illustrates the law of definite proportions.

6. 11.7gms. of a sample of sodium chloride was found to contain 4.6 gms. of sodium and 7.1 gms. of chlorine in it. 2.93 gms. of an another sample of sodium chloride was found to contain 1.15 gms. of sodium and 1.78 gms. of chlorine. Show that these figures illustrate the law of definite proportion.

Both the samples contain the same elements sodium and chlorine. If they were to illustrate the law of constant proportion, the percentage of sodium and chlorine in the above two samples should remain the same.

In the first sample, 11.7 gms. of the sample has 4.6 gms. of sodium.

$$\therefore \text{Percentage of sodium} = \frac{\text{Weight of sodium}}{\text{Weight of sample}} \times 100$$

$$= \frac{4.6}{11.7} \times 100 = 39.31\%$$

In the second sample, 2.93 gms. of the sample has 1.15 gms. of sodium.

$$\therefore \text{Percentage of sodium} = \frac{\text{Weight of sodium}}{\text{Weight of sample}} \times 100$$

$$= \frac{1.15}{2.93} \times 100 = 39.31\%$$

Since the two samples contain the same elements combines together in the same proportion by weight, the law of definite proportion is proved.

7. In an experiment 10 gms of calcium combines with 4 gms of oxygen to form the oxide of calcium. In another experiment 15 gms of calcium combines with 6 gms of oxygen to form the oxide of calcium. Show that this data illustrates the Law of Definite Proportions.

Solution

	Expt I	**Expt II**
Wt. of Calcium	10 gms	15 gms
Wt. of Oxygen	4 gms	6 gms
Wt. of Calcium: Wt. of Oxygen	10: 4 :5:2	15:6 : 5:2

In both cases, the ratio of the weight of calcium to he weight of oxygen is a constant. i.e., 5:2. Thus these results illustrate the Law of Definite Proportions.

8. 0.2432 gm of magnesium when burnt in air yielded 0.4032 gm of magnesium oxide while 0.1824 gm of magnesium gave 0.3024 gm of magnesium oxide. Show that these results are in accordance with the Law of Definite Proportions.

Solution

	Expt I	**Expt II**
Wt. of Magnesium oxide	0.4032 gms	0.3024 gms
Wt. of Magnesium	0.2432 gms	0.1824 gms
Wt. of Oxygen	0.1600 gms	0.1200 gms
Wt. of Magnesium: Wt. of Oxygen	0.2432: 0.1600 1: 52 :1	0.1824: 0.1200 1:52 : 1

Thus the ratio of the weight of magnesium to the weight of oxygen is found to be the same in both cases. This is in accordance with the Law of Definite Proportions.

9. In one experiment, 0.5 g of copper was converted into cupric oxide and its weight is found to be 0.6248 g. In a second experiment, 1.25 g of cupric oxide is obtained from 1.00 g of copper by the same method. Show that these results illustrate the law of definite proportions.

	Expt I	**Expt II**
Wt. of Cupric oxide	0.6248	1.25
Wt. of Copper	0.5000	1.00
Wt. of oxygen	0.1248	0.25
Wt. of Copper: Wt. of Oxygen	0.5000: 0.1248 4:1	1.00: 0.25 4:1

These results illustrate the law of definite proportions i.e., a compound by whatever method it is prepared always contains the same elements in the same fixed proportions by weight.

10. In two experiments, 0.259 g and 0.207 g of lead were converted to lead chloride. 0.347 g and 0.278 g of lead chlorine are formed. Show that this data illustrate the law of constant proportions.

Solution. In the first experiment,

Weight of lead chloride = 0.347 g

Weight of lead = 0.259 g

$\therefore$ Weight of chlorine = 0.347 – 0.259

= 0.088 g

The ratio of lead and chlorine = 0.259: 0.088

= 3:1

In the second experiment,

Weight of lead chloride = 0.278 g

Weight of lead = 0.207 g

∴ Weight of chlorine = 0.278 – 0.207

= 0.071 g

The ratio of lead and chlorine = 0.207: 0.071

= 3:1

Since in both experiments, it is found that in lead chlorine, the lead and chlorine are present in the ratio 3:1. This proves the law of constant composition.

11. 1.375 g of cupric oxide on reduction in hydrogen gas gives 1.098 g of copper. In another experiment, 1.179 g of metallic copper produced 1.476 g of copper oxide. Show that these results illustrate the law of constant (or definite) proportions.

Solution.

Experiment 1: Mass of copper oxide taken = 1.375 g

Mass of copper obtained = 1.098 g

Therefore Mass % of copper $= \dfrac{1.098 \times 100}{1.375} = 79.86$

Experiment 2: Mass of copper oxide produced = 1.476 g

Mass of copper used = 1.179 g

Therefore Mass % of copper $= \dfrac{1.179 \times 100}{1.476} = 79.89$

Since, the percentage of copper in the two samples of copper oxide is the same, hence the law of definite proportion is verified.

12. 1.19 g of zinc was converted into zinc oxide and 1.51 gm of zinc oxide was obtained. In another experiment 1.812 gm of zinc oxide when heated with carbon gave 1.428 gm of zinc. Show how these results illustrate the law of definite proportions.

	Expt I	**Expt II**
Wt. of zinc oxide	1.51 gm	1.812 gm
Wt. of zinc	1.19 gm	1.428 gm
Wt. of oxygen	0.32 gm	0.384 gm
Wt. of zinc: Wt of oxygen	**1.19: 0.32 = 3.72: 1**	**1.428: 0.384 = 3.72: 1**

Thus the ratio of weight of zinc to the weight of oxygen is found to be the same in both the cases. This illustrates the law of definite proportions.

13. In one experiment 1.098 g of copper is obtained by the reduction of 1.375 g of cupric oxide. In another experiment 1.476 g of cupric oxide was prepared from 1.179 g of copper through cupric nitrate. Show that the results of the two experiments illustrate the law of definite proportions?

Experiment: 1

Weight of cupric oxide = 1.375 g

Weight of copper = 1.098 g

Weight of oxygen = 0.277 g

Ratio in the weights of copper and oxygen = 1.098: 0.277 = **4:1**

Experiment: 2

Weight of the cupric oxide = 1.476 g

Weight of the copper = 1.179 g

Weight of oxygen = 0.297 g

Ratio in the weights of copper and oxygen = 1.179: 0.297

= **4:1**

In both cases, it is found that copper and oxygen are in the ratio **4:1,**
Hence, the law of definite proportion is proved.

14. Illustrate the law of definite proportions from the following data:

(i) 0.16 g of sulphur produces 0.32 g of sulphur dioxide, (ii) sulphur dioxide obtain by the decomposition of sodium sulphate contains 50% sulphur.

Case 1

(i) Weight of sulphur dioxide	=	0.32 g
Weight of sulphur	=	0.16 g
Weight of oxygen	=	0.16 g

Ratio of sulphur to oxygen in sulphur dioxide produced by (i) Method = **0.16: 0.16 or 1:1**

Case 2

(ii) Weight of sulphur dioxide	=	100 g
Weight of sulphur	=	50 g
Weight of oxygen	=	50 g

Ratio of sulphur to oxygen in sulphur dioxide produced by (ii) Method = **50:50 or 1:1**

Since the sulphur dioxide produced by different methods contains sulphur and oxygen in the ratio 1:1, the data proves **the law of definite proportions.**

15. In two experiments, 0.259 g and 0.207 g of lead were converted to lead chloride, yielding 0.347 g and 0.278 g of lead chlorided respectively. Show that the data illustrates the law of constant composition.

Experiment: 1

Weight of lead chloride	=	0.347 g
Weight of lead	=	0.259 g
Weight of chlorine	=	0.088 g

In this experiment, the lead and chlorine are present in the ratio **0.259: 0.088 or 3:1.**

In the second experiment

Experiment: 2

Weight of the lead chloride	=	0.278 g
Weight of the lead	=	0.207 g
Weight of chlorine	=	0.071 g

The lead and chlorine are present in the ratio **0.207: 0.071 or 3:1.**

Since in both the experiments, it is found that in lead chloride, lead and chlorine are present in the ratio 3:1. This proves **the law of constant composition.**

III. *LAW OF MULTIPLE PROPORTION*:

1. Two oxides of metal contain 20% and 11.1% oxygen respectively. Show how these data illustrate law of multiple proportions

Case 1

Weight of the first oxide	=	100 g
Weight of oxygen	=	20 g
$\therefore$ Weight of the metal	=	80 g

80 g of metal combines with 20 g of oxygen

1 g of metal combines with $\frac{20}{80} \times 1$ of oxygen

= **0.25g of oxygen ……….. (1)**

Thus, the ratio of the metal to oxygen in the first oxide is 1: 0.25.

Case 2

Weight of the second oxide = 100.0 g
Weight of oxygen = 11.1 g

Weight of the metal = 88.9 g

88.9 g of metal combines with = 11.1 g of oxygen

$\therefore$ 1 g of metal combines with = $\frac{11.1}{88.9}$ g of oxygen

= **0.12 g of oxygen ………… (2)**

Thus, the ratio of the metal to oxygen in the second oxide is 1: 0.12.

From (1) & (2) the different weights of oxygen that combines with fixed weigh of the metal (i.e., 1.0 g) are in the ratio **0.25: 0.12** or **2:1.** This is a simple ratio and hence the law of multiple proportions is proved.

2. A metal forms three oxides containing respectively 76.47%, 68.42% and 52.2% of the metal. Show that these data are in accordance with law of multiple proportions.

Case 1

Weight of metallic oxide = 100.00 g
Weight of the metal = 76.47 g

Weight of oxygen = 23.53 g

23.53 g of oxygen combines with 76.47 g metal

1 g of oxygen will combine with $\frac{76.47}{23.53}$ = **3.2 g of metal ……….. (1)**

Case 2

Weight of the metallic oxide = 100.00 g
Weight of the metal = 68.42 g

Weight of oxygen = 31.58 g

31.53 g of oxygen combines ith 68.42 g of the metal

1 g of oxygen will combine with $\frac{68.42}{31.53}$ = **2.1 g of metal ……… (2)**

Case 3

Weight of the metallic oxide = 100.0 g
Weight of the metal = 52.2 g

Weight of oxygen = 47.8 g

47.8 g oxygen combines with 52.2 g of metal

1 g of oxygen will combine with $\frac{52.2}{47.8}$ = **1.09 g of metal ……….. (3)**

In all these cases, different weights of the metal combine with the fixed weight of oxygen (i.e., 1g). According to the law of multiple proportions, they are in a simple integral ratio. i.e., 3:2:1. Hence, these data are in accordance with the **law of multiple proportions**.

3. Carbon and oxygen are known to form two compounds. The carbon content in one of these is 42.9% while in the other it is 27.3%. Show that this data is in agreement with the law of multiple proportions.

Solution. For first compound

Mass % of C = 42.9

$\therefore$ Mass % of O = 57.1

Thus, 42.9 g of C reacts with 57.1 g of oxygen

1 g of C reacts with 57.1 / 42.9 g oxygen = 1.33 g of oxygen

For second compound

	Mass % of C	=	27.3
∴	Mass % of O	=	72.7

Thus, 27.3 g of C reacts with 72.7 g of oxygen

1 g of C reacts with 72.7 / 27.3 g of oxygen = 2.66 g of oxygen

The ratio of oxygen masses which combine with 1 g of C is

1.33 : 2.66 or 1: 2

Since, 1: 2 is a simple ratio, hence the law of multiple proportions is supported by the data.

4. In one experiment 0.75 g of metal was converted into its oxide which weighed 0.9389 g and in a second experiment 1 g of metallic oxide on reduction gave 0.7989 g of metal. Show how these results illustrate a Law of Chemical Combinations.

Let us first find out the weights in the two oxides.

In one, 0.75 g of Metal combines with (0.9389 – 0.75) g

= 0.1889 g of oxygen

In the other, 0.7989 g of Metal combines with (1- 0.7989) g

= 0.3011 g of oxygen.

Now, fixing the weight of Metal, weights of oxygen are calculated.

In one, ∴0.75 g of Metal combines with 0.1889 g of oxygen

$$\therefore 1 \text{ g of Metal combines with} = \frac{0.1889 \times 100}{75} = 0.252 \text{ g}$$

In the other

∴0.7989 g of Metal combines with 0.3011 g of oxygen

$$\therefore 1 \text{ g of Metal combines with } \frac{0.3011}{0.7989} = 0.376 \text{ g}$$

The two weights of oxygen combining with a fixed weight of copper are

0.252 g and 0.376 g

The two weights are in the ratio

$$\frac{0.252}{0.376} = 2 : 3$$

The results illustrate the Law of **Multiple Proportion**. The Law is stated as:

When A and B combine to form two or more than two compounds, the different weight of one combining with a fixed weight of B bear a simple ratio.

5. Two oxides of a metal contained respectively 7.41% and 3.85% of oxygen. Show that these facts agree with the Law of Multiple Proportions.

Solution

	Expt I	**Expt II**
Wt. of oxygen	7.41 gms	3.85 gms
Wt. of metal	92.59 gms	96.15 gms
Wt. of Oxygen in combination with 1 gm of metal	7.41 / 92.59 gms 0.08 gm	3.85 / 96.15 gms 0.04 gm

Thus the different weights of oxygen that combine with the same weight of the metal, namely 1 gm, are 0.08 gm and 0.04 gm. The ratio of wts. of Oxygen = 0.08: 0.04

= 2 : 1

which is a simple integral ratio. Thus the Law of Multiple Proportions is illustrated.

6. 1.90 gm of one oxide of copper gave 1.52 gm of copper on reduction. 2.85 gm of another oxide gave 2.53 gm of copper on reduction. Show that these results are in accordance with the Law of Multiple Proportions.

Solution

	Expt I	**Expt II**

Wt. of Copper oxide	1.90 gm	2.85 gm
Wt. of Copper	1.52 gm	2.53 gm
Wt. of oxygen	0.38 gm	0.32 gm
Wt. of Oxygen combining with 1 gm of Copper	0.38 / 1.52 = 0.25 gm	0.32 / 2.53 = 0.13 gms

The ratio of the wts. of Oxygen = 0.25: 0.13
= 2:1 (appr). which is a simple integral ratio
Thus the given results are in accordance with the Law of Multiple Proportions.

7. 3 gms of carbon form two types of oxides with weights as 7 and 11 gms respectively. Show that these illustrate law of multiple proportions:
Solution:

1st oxide
Weight of oxide = 7 gms.
Weight of carbon = 3 gms.
Weight of oxygen = 4 gms.
2nd oxide
Weight of oxide = 11 gms.
Weight of carbon = 3 gms.
Weight of oxygen = 8 gms.

The weight of oxygen that combines with a fixed weight of carbon, namely 3 gms of carbon are 4 gms. and 8 gms. respectively.

$\therefore$ The ratio of the weight of oxygen = 4:8 = 1:2

The ratio is a simple integral ratio. Thus the given data illustrates the law of multiple proportions.

8. In a typical experiment, 4 gms. of hydrogen are found to form 36 gms. of water. In an another experiment 2 gms. of hydrogen are found to form 34 gms. of hydrogen peroxide. Show that these illustrate the law of multiple proportions.
Solution:
Case 1: Weight of water = 36 gms.
Weight of hydrogen = 4 gms.

Weight of oxygen = 32 gms.
----

4 gms. of hydrogen combine with = 32 gms. of oxygen.

$\therefore$ 1 gm. of hydrogen combines with $= \frac{32 \times 1}{4} = 8$ gms.

Case 2: Weight of hydrogen peroxide = 34 gms.
Weight of hydrogen = 2 gms.

Weight of oxygen = 32 gms.
----

2 gms. of hydrogen combine with = 32 gms. of oxygen.

$\therefore$ 1 gm. of hydrogen combines with $= \frac{32 \times 1}{2} = 16$ gms. of oxygen.

$\therefore$ The different weights of oxygen that combine with the same weight of hydrogen, namely 1 gm. of hydrogen are 8 and 16 gms.

The ratio of oxygen weights = 8 : 16
= 1 : 2, a simple integral ratio. Thus the data illustrates the law of multiple proportions.

9. In an experiment 2.8 gm of nitrogen gave 6 gm of its oxide. In another experiment 2.1 gm of nitrogen gave 6.9 gm of another oxide. Show how these results illustrate the law of multiple proportions.

	Expt I	**Expt II**
Wt. of nitrogen oxide	6.0 gm	6.9 gm

Wt. of nitrogen Wt. of oxygen Wt. of oxygen combining with 1 gm of nitrogen	2.8 gm 3.2 gm 3.2 / 2.8 = 1.142 gm	2.1 gm 4.8 gm 4.8 / 2.1 = 2.285 gms

Thus the different weight of oxygen that combines with the same weight of nitrogen (1 g), are 1.142 g and 2.285 g. i.e., **the ratio of weights of oxygen = 1:2**.which is a simple integral ratio. Thus the law of multiple proportions is illustrated.

10. Two oxides of metal contain 20% and 11.1% oxygen respectively. Show how these data illustrate law of multiple proportions

Case 1

Weight of the first oxide = 100 g

Weight of oxygen = 20 g

$\therefore$ Weight of the metal = 80 g

----

80 g of metal combines with 20 g of oxygen

1 g of metal combines with $\frac{20}{80}$ x 1 of oxyge

= **0.25g of oxygen ……….. (1)**

Thus, the ratio of the metal to oxygen in the first oxide is 1: 0.25.

Case 2

Weight of the second oxide = 100.0 g

Weight of oxygen = 11.1 g

Weight of the metal = 88.9 g

88.9 g of metal combines with = 11.1 g of oxygen

$\therefore$ 1 g of metal combines with = $\frac{11.1}{88.9}$ g of oxygen

= **0.12 g of oxygen ………… (2)**

Thus, the ratio of the metal to oxygen in the second oxide is 1: 0.12.

From (1) & (2) the different weights of oxygen that combines with fixed weigh of the metal (i.e., 1.0 g) are in the ratio **0.25: 0.12** or **2:1.** This is a simple ratio and hence the law of multiple proportions is proved.

11. Copper combines with oxygen to form two oxides, which have the following composition:

(i) 0.716 g of cuprous oxides contains 0.630 g of copper.

(ii) 0.398 g of cupric oxide contains 3.318 g of copper.

Prove that the above data illustrates the law of multiple proportions.

Solution: (i) In the first experiment,

Weight of cuprous oxide = 0.716 g

Weight of copper = 0.630 g

$\therefore$ Weight of oxygen = 0.716 – 0.630

= 0.086 g

(ii) In the second experiment,

Weight of cupric oxide = 0.398 g

Weight of copper = 0.318 g

$\therefore$ Weight of oxygen = 0.398 – 0.318

= 0.08 g

Here copper is forming two oxides. The weight of oxygen in both cases is the same i.e., **0.08 g**. Thus, a definite weight of oxygen combines with 0.630 g and 0.318 g of copper. The proportion of weight of copper in these compounds

Cuprous oxide	:	cupric oxide
0.630 g	:	0.318 g

2 : 1

The proportion by weight of copper is indicated by a simple ratio. Thus, the law of multiple proportions is obeyed.

12. A metal forms three oxides containing respectively 76.47%, 68.42% and 52.2% of the metal. Show that these data are in accordance with law of multiple proportions.

Case 1

Weight of metallic oxide = 100.00 g
Weight of the metal = 76.47 g

Weight of oxygen = 23.53 g

23.53 g of oxygen combines with 76.47 g metal

$$1 \text{ g of oxygen will combine with } \frac{76.47}{23.53} = \textbf{3.2 g of metal ……….. (1)}$$

Case 2

Weight of the metallic oxide = 100.00 g
Weight of the metal = 68.42 g

Weight of oxygen = 31.58 g

31.53 g of oxygen combines ith 68.42 g of the metal

$$1 \text{ g of oxygen will combine with } \frac{68.42}{31.53} = \textbf{2.1 g of metal ……… (2)}$$

Case 3

Weight of the metallic oxide = 100.0 g
Weight of the metal = 52.2 g

Weight of oxygen = 47.8 g

47.8 g oxygen combines with 52.2 g of metal

$$1 \text{ g of oxygen will combine with } \frac{52.2}{47.8} = \textbf{1.09 g of metal ……….. (3)}$$

In all these cases, different weights of the metal combine with the fixed weight of oxygen (i.e., 1g). According to the law of multiple proportions, they are in a simple integral ratio. i.e., 3:2:1. Hence, these data are in accordance with the **law of multiple proportions**.

IV. *LAW OF RECIPROCAL PROPORTIONS*:

1. Water and sulphur dioxide contains 88.9% and 50% oxygen respectively. Hydrogen sulphide contains 91.1% of sulphur. Illustrate the law of reciprocal proportions from these data?

Case 1

Weight of water = 100g
Weight of oxygen = 88.9 g

Weight of hydrogen = 11.1 g

11.1 g of hydrogen combine with 88.9 g oxygen.

Let the weight of hydrogen be fixed as one gram.

$$\text{Hence, 1.0 g hydrogen will combine with } \frac{88.9}{11.1} = \textbf{8.0 g} \text{ of oxygen.}$$

In water hydrogen and oxygen are present in the ratio **1: 8……… (1)**

Case 2

Similarly in hydrogen sulphide

Weight of hydrogen sulphide = 100.0g
Weight of sulphur = 91.1g

Weight of hydrogen = 8.9g

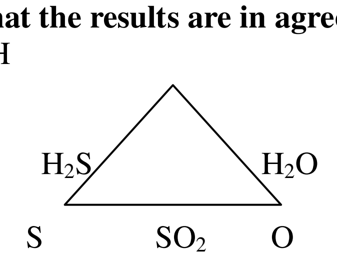
8.9 g of hydrogen combines with 91.1 g of sulphur. 1 g of hydrogen will combine with

$$\frac{91.1}{8.9} = 10.2 \text{ g of sulphur.}$$

In hydrogen sulphide, hydrogen and sulphur are present in the ratio **1: 10.2…. (2)**

The ratio of sulphur to oxygen is 1: 1 in sulphur dioxide. From (1) & (2) , it can be shown that sulphur and oxygen must be present in the ratio 10.2: 8 or 1: 0.78 or 1:1, Hence, the law of reciprocal proportions is verified.

2. Hydrogen sulphide contains 94.11% Sulphur. Sulphur dioxide contains 50% oxygen. Water contains 11.11% hydrogen. Show that the results are in agreement with the law of reciprocal proportions.



Solution. (a) Calculation of masses of Sulphur and oxygen that combine with certain fixed mass, say 1 g of hydrogen.

(i) In H_2S, 100 – 94.11 = 5.89 g hydrogen combines with 94.11 g Sulphur. So,

$$1 \text{ g hydrogen combines with } \frac{94.11}{5.89} \text{ g} = 15.98 \text{ g of Sulphur}$$

(ii) In water, 11.11 g hydrogen combines with 100 – 11.11 = 88.89 g oxygen. So,

$$1 \text{ g hydrogen combines with } \frac{88.89}{11.11} \text{ g} = 8 \text{ g of oxygen}$$

(b) Calculation of ratio of the masses of Sulphur (in H_2S) and oxygen (in H_2O)

$$\frac{\text{Mass of Sulphur}}{\text{Mass of oxygen}} = \frac{15.98}{8} = 2$$

(c) Calculation of the ratio of masses of Sulphur and oxygen when they combine to form Sulphur dioxide.

$$\frac{\text{Mass of Sulphur}}{\text{Mass of oxygen}} = \frac{50}{50} = 1$$

The ratio (i) is double of (ii), i.e., 2:1, which is a simple ratio. This illustrates the law of reciprocal proportions.

3. Phosphorus trioxide contains 56.4% of phosphorus and 43.6% of oxygen. Water contains 88.8% of oxygen and 11.2% of hydrogen. Phosphine contains 91.1% of phosphorus and 8.9% of hydrogen. Show how these results illustrate the Law of Reciprocal Proportion.

	Phosphorus	**Oxygen**	**Hydrogen**
P_2O_3	56.4%	43.6%	-
H_2O	-	88.8%	11.2%
PH_3	91.1%	-	8.9%

First the ratio in the weight of phosphorus and hydrogen which combine with a fixed weight of oxygen is calculated.

$\therefore$45.6 g of oxygen combines with 56.4 g of phosphorus

$\therefore$1 g of oxygen combines with 56.4 / 43.6 g of phosphorus

$\therefore$88.8 g of oxygen combines with 11.2 g of hydrogen

$\therefore$1 g of oxygen combines with 11.2 / 88.8 g of hydrogen

The ratio of the weights of phosphorus and hydrogen combining with 1 g of oxygen in P_2O_3 and H_2O is

56.4 11.2

$$= \frac{\quad}{43.6} / \frac{\quad}{88.8} = 1.29 / 0.12 = 10.2$$

And the ratio of weights of phosphorus and hydrogen in PH_3 is 91.1 /8.9 = 10.2. Thus the two ratios are related as 10.2: 10.2 = 1:1.

4. Methane contains 75% of carbon and 25% of hydrogen. Carbon monoxide contains 42.86% of carbon and 57.14% of oxygen. Water contains 11.11% of hydrogen and 88.89% of oxygen. What law does it illustrate?

Compound	Weight of Carbon	Weight of Hydrogen	Weight of Oxygen
Methane	75%	25%	-
Carbon monoxide	42.86%	-	57.14%
Water	-	11.11%	88.89%

The ratio by weight in which hydrogen and oxygen combine among themselves = 11.11 / 88.89 = 0.1251 …..1

Let the fixed weight of carbon be 75 g.

Weight of oxygen combining with 42.86 g of carbon in carbon monoxide = 57.14 g.

Weight of oxygen combining with 75 g of carbon = $\frac{75 \times 57.14}{42.86}$

= 100 g

Weight of hydrogen combining with 75 g of carbon = 25.0 g

∴ Ratio of weights of hydrogen and oxygen, which combine with carbon,

25 / 100 = 0.250 …..2

The ratio between the two values (1) & (2) = 0.125 : 0.250

= 1:2

Hence the law of reciprocal proportions is obeyed.

5. Carbon dioxide contains 27.27% carbon, carbon disulphide contains 15.79% carbon and Sulphur dioxide contains 67% Sulphur. Show that these data are in accordance with the law of reciprocal proportions.

Case 1 Weight of carbon dioxide = 100.00 g
Weight of carbon = 27.27 g

Weight of oxygen = 72.73 g

27.27 g of carbon combines with 72.73 g of oxygen.

∴ 1g of carbon combines with $\frac{72.73}{27.27}$ = 2.66 g of oxygen.

Ratio of carbon and oxygen in carbon dioxide = 1:2.6

Case 2 Weight of carbon disulphide = 100.00 g
Weight of carbon = 15.79 g

Weight of Sulphur = 84.21 g

∴ 15.79 g of carbon combines with 84.21 g of Sulphur

1 g of carbon combines with $\frac{84.21}{15.79}$ = 5.3 g of Sulphur.

∴ Ratio of carbon and Sulphur in carbon dioxide = 1: 5.3

If Sulphur and oxygen were to combine to sulphur dioxide, according to the law of reciprocal proportions, they must combine in the ratio 5.3: 2.6 (or) 2:1.

Case 3 Weight of sulphurdioxide = 100 g
Weight of Sulphur = 67 g

Weight of oxygen = 33 g

In the case of SO_2 67 g of Sulphur combines with 33g of oxygen.
$\therefore$ The ratio Sulphur to oxygen is 67:33 (or) 2.1:1
Hence, these data are in accordance with law of reciprocal proportions.

6. Potassium chloride contains 52% potassium, potassium iodide contains 23.6% potassium and iodine chloride contains 78.2% iodine. Show that these illustrate law of reciprocal proportions.

Case 1 Weight of potassium chloride = 100 g
Weight of potassium = 52 g

Weight of chlorine = 48 g

----

52 g of potassium combines with = 48 g of chlorine
1 g of potassium combine with = 48 / 52 = 0.92 **g** of chlorine. ..1

Case 2 Weight of potassium iodide = 100.0 g
Weight of potassium = 23.6 g

Weight of iodine = 76.4 g

23.6 g of potassium combines with = 76.4 g iodine.
1 g of potassium combines with = 76.4 / 23.6 = 3.2 g of iodine ..2

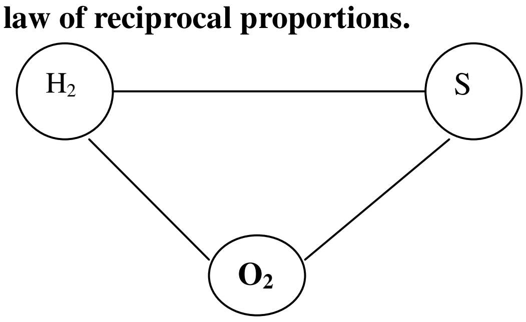
According to the law of reciprocal proportions, if iodine and chlorine were to combine, they will combine in the ratio **3.2: 0.92 (or) 7:2**.

Case 3 Weight of iodine chloride = 100.0 g
Weight of iodine = 78.2 g

Weight of chlorine = 21.8 g ….3

78.2 g iodine combines with 21.8 g chlorine.
The ratio of iodine and chlorine in iodine chloride **78.2:21.8 (or) 3.5:1 (or) 7:2**.
Hence, the results prove **the law of reciprocal proportions**.

7. One gram of hydrogen combines with 15.88 g of Sulphur. One gram of hydrogen combines with 7.92 of oxygen, one gram of Sulphur combines with 0.998 g of oxygen. Show that these data illustrate the law of reciprocal proportions.



	Weight of H_2	Weight of O_2	Weight of S
H_2S	1 g	-	15 g
H_2O	1 g	7.92 g	-
SO_2	-	0.998 g	1 g

In H_2S

Weight of H_2 = 1 g
Weight of S = 15 g

In H_2O

Weight of H_2 = 1 g
Weight of O_2 = 7.92 g
$\therefore$ The ratio between S in H_2S and O_2 in water is 15:7.98 i.e., 2:1 1

In SO_2

Weight of S = 1 g
Weight of O_2 = 0.998 g
$\therefore$ The ratio between S and O_2 in SO_2 is 1:0.998 i.e., 1:1 2

The two ratios (1) and (2) related as 2 / 1: 1 / 1 or 2:1
$\therefore$ The law of reciprocal proportions holds good.

8. In an experiment 34.5 g oxide of a metal was heated so that O_2 was liberated and 32.1 g of metal was obtained. In another experiment 119.5 g of another oxide of the same metal was heated and 103.9 g metal was obtained and O_2 was liberated. Calculate the mass of O_2 liberated in each experiment. Show that the data explain the law of multiple proportions.

Solution: In the first experiment

Weight of the metal oxide	=	34.5 g
Weight of the metal	=	32.1 g
Weight of oxygen liberated	=	2.4 g

32.1 g metal combines with 2.4 g oxygen.

1 g of the metal combines with 2.4 / 32.1 = 0.075 g

In the second experiment,

Weight of the oxide taken	=	119.5 g
Weight of the metal formed	=	103.9 g
Weight of oxygen liberated	=	15.6 g

103.9 g of metal combines with 55.6 g oxygen.

$$1 \text{ g of metal} = \frac{15.6}{103.9} \times 1 = 0.15014 \text{ oxygen}$$

Therefore different weights of oxygen, that combine with the fixed weight of the metal viz 1 g are in the ratio

0.15014 : 0.075
2 : 1

9. Water and sulphur dioxide contains 88.9% and 50% oxygen respectively. Hydrogen sulphide contains 91.1% of sulphur. Illustrate the law of reciprocal proportions from these data?

Case 1

Weight of water	=	100g
Weight of oxygen	=	88.9 g
Weight of hydrogen	=	11.1 g

11.1 g of hydrogen combine with 88.9 g oxygen.
Let the weight of hydrogen be fixed as one gram.

Hence, 1.0 g hydrogen will combine with $\frac{88.9}{11.1}$ = **8.0 g** of oxygen.

In water hydrogen and oxygen are present in the ratio **1: 8.......... (1)**

Case 2

Similarly in hydrogen sulphide

Weight of hydrogen sulphide	=	100.0g
Weight of sulphur	=	91.1g
Weight of hydrogen	=	8.9g

8.9 g of hydrogen combines with 91.1 g of sulphur. 1 g of hydrogen will combine with
$\frac{91.1}{}$ = 10.2 g of sulphur.

8.9

In hydrogen sulphide, hydrogen and sulphur are present in the ratio **1: 10.2…. (2)**

The ratio of sulphur to oxygen is 1: 1 in sulphur dioxide. From (1) & (2) , it can be shown that sulphur and oxygen must be present in the ratio 10.2: 8 or 1: 0.78 or 1:1, Hence, the law of reciprocal proportions is verified.

4 – CHEMICAL REACTIONS & EQUILIBRIUM

INTRODUCTION:

In a chemical reaction atoms or groups of atoms carrying electrical charge (called radicals) are transferred from one substance to another and new substances are formed from the old ones. A chemist is interested in the following aspects of a chemical reaction:

(i) The rate with which the reaction occurs and factors affecting the rate and

(ii) Mechanism of the reaction.

Many reactions go to completion. However, there are some which never go to completion. Such reactions proceed upto a certain point at which they apparently seem to stop. The reaction mixture at this point contains products as well as unconsumed reactants. For such a reaction, the ratio of product of concentrations of products and that of reactants has a definite value at a given temperature. When the reaction reaches that point, a state of equilibrium is reached. The equilibrium so established is dynamic in nature and can be shifted by alterations in temperature, pressure and concentration of reactants and products.

IMPORTANT TERMS & DEFINITIONS

Instantaneous (or) fast reactions: These reaction occur at once, for example, ionic reactions such as acid-base neutralization reaction and precipitation reaction $NaOH + HCl \longrightarrow NaCl + H_2O$

Reactions of this type proceed so quickly and their rates cannot be determined by common methods.

Rate of reaction: Rate or velocity of a chemical reaction is defined as "change in molar concentration of one of the reactants or products in unit time".

$$\text{Rate} = \frac{dx}{dt} = \frac{\text{Change in molar concentration}}{\text{Time}}$$

$$\text{Unit: Rate} = \frac{dx}{dt} = \frac{\text{mol / dm}^3}{\text{second}}$$

Rate of the reaction is expressed in mol /dm^3 /s or mol dm^{-3} s^{-1}.

Law of mass action: At constant temperature, the rate of a chemical reaction is directly proportional to the product of 'active masses' of the reactants.

$$\text{Atomic mass} = \frac{\text{No of moles}}{\text{Volume of the containing vessel}}$$

Photochemical reaction: A few chemical reactions take place only when the reacting substances after mixing is exposed to light.

Chemical equilibrium: When the rate of forward reaction becomes equal to the rate of backward reaction. This state is called chemical equilibrium.

Thermo chemistry is a branch of chemistry which deals with the study of heat change during various physical and chemical transformations.

Endothermic reaction: A reaction which takes place with the absorption of heat is called endothermic reaction.

For an endothermic reaction, $H_p > H_R$. Hence $\Delta H = +ve$.

'Active mass' means effective molar concentration and in a dilute solution, it can be considered to be equal to molar concentration expressed in mol/dm^3 or partial pressure expressed in (N/m^2) in the case of gases.

Exothermic reaction: A reaction which takes place with the evolution of heat is known as exothermic reaction. For an exothermic reaction, $H_p < H_R$. Hence, $\Delta H = -$ ve. *eg. (i)* $C(s) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = -393.5$ kJ

Reasons for difference in rates: A chemical reaction involves the breaking and making of bonds. A strong bond requires more energy to break, than a weak bond. A reaction which involves breaking of strong bonds, will be slow or impossible to occur at room temperature. On the other hand, a reaction which involves breaking of only weak bonds will be fast under similar conditions.

Irreversible reaction: A reaction in which the products formed do not recombine to produce the original reactants is called an irreversible reaction.

Reactions involving liberation of a gas or those in which a precipitate is formed are generally **irreversible reactions**. *eg.:* $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$

Conditions for reversible reaction: 1. The reaction should be done in a closed vessel. 2.None of the products should be removed from the vessel. 3. Temperature and pressure should be kept constant.

Reversible reaction: A reaction in which the products formed react to give back the original substances is called a reversible reaction. *Example:* $N_2 + 3H_2 \longrightarrow 2NH_3$. In a reversible reaction, the reaction proceeding from left to right is called a forward reaction. The reaction proceeding from right to left is called a backward reaction or reverse reaction.

Factors influencing the rate of a reaction: 1. Concentration of the reactants 2. Temperature: 3. Presence of a Catalyst 4. Nature of reactants 5. Nature of the solvent 6. Exposure to radiations 7. Surface area

Classification of the reactions:

(a) Instantaneous (or) fast reactions: These reaction occur at once, for example, ionic reactions such as acid-base neutralization reaction and precipitation reaction $NaOH + HCl \longrightarrow NaCl + H_2O$

Reactions of this type proceed so quickly and their rates cannot be determined by common methods.

(b) Extremely slow reactions: Some reactions proceed at extremely slow rates. *For example*, rusting of iron, combination of hydrogen and oxygen at room temperature *etc.,* are few such reactions, which takes months, or even years before any observable change occur.

(c) Reactions with moderate speeds: In between the two extremes discussed above, there are a large number of reactions, which proceeds at moderate rates. A few examples of this type are given below.

(i) Decomposition of hydrogen peroxide $2H_2O_{2\,(liq)} \longrightarrow 2H_2O_{(liq)} + O_2(g)$

SELF EVALUATION (T.B.PAGE 74)

I. Choose the correct answer.

1. Rusting of iron is an example for ....
   (a) Fast reaction (b) slow reaction
   (c) moderate reaction (d) photochemical reaction
2. For an exothermic reaction ....
   (a) $H_P < H_R$ (b) $H_R = H_P$
   (c) $H_p > H_R$ (d) $H_p \neq H_R$
3. Unit for the rate of the reaction is ....
   (a) mol $dm^{-3}s^{-1}$ (b) mol dm^3s^{-1}
   (c) mol dm^3s (d) mol^{-1} $dm^{-3}s^{-3}$
4. Ionic reaction is an example for ....
   (a) instantaneous reaction (b) slow reaction
   (c) molecular reaction (d) photochemical reaction
5. $2H_2O_2(l) \longrightarrow 2H_2O(l) + O_2(g)$. This reaction is an example for .... reaction.
   (a) Instantaneous reaction (b) slow reaction
   (c) molecular reaction (d) photochemical reaction
6. Dissolution of sodium carbonate in H_2O is an example for .... reaction.
   (a) Endothermic (b) exothermic
   (c) photochemical (d) electrochemical
7. The value of ΔH is positive for an ....
   (a) Endothermic (b) exothermic
   (c) photochemical (d) electrochemical
8. The rate of the reaction .... with increase in concentration.
   (a) increases (b) decreases
   (c) remains constant (d) becomes zero
9. Active mass of substance expressed in ....
   (a) mol dm^3 (b) mol dm^{-3}
   (c) mol^{-1} dm^{-3} (d) mol^{-3} dm
10. The speed of the reaction .... for every 10°C rise in temperature.
    (a) doubles (b) increases
    (c) decreases (d) remains constant

Answers:

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

II. Answer the following in One or Two sentences. (Page 75)

1. **What are instantaneous reactions?**

These reaction occur at once, for example, ionic reactions such as acid-base neutralization reaction and precipitation reaction

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

$$AgNO_3 + KCl \longrightarrow AgCl\downarrow + KNO_3$$

Reactions of this type proceed so quickly and their rates cannot be determined by common methods.

2.Define rate of reaction.

Rate or velocity of a chemical reaction is defined as "change in molar concentration of one of the reactants or products in unit time".

$$\text{Rate} = \frac{dx}{dt} = \frac{\text{Change in molar concentration}}{\text{Time}}$$

$$\text{Unit: Rate} = \frac{dx}{dt} = \frac{\text{mol / dm}^3}{\text{second}}$$

Rate of the reaction is expressed in mol /dm^3 /s or mol dm^{-3} s^{-1}.

3.Define law of mass action.

At constant temperature, the rate of a chemical reaction is directly proportional to the product of 'active masses' of the reactants.

$$\text{Atomic mass} = \frac{\text{No of moles}}{\text{Volume of the containing vessel}}$$

4.Define photochemical reaction.

Photochemical reaction: A few chemical reactions take place only when the reacting substances after mixing is exposed to light.

Example: The reaction between hydrogen and chlorine in the presence of sunlight to give hydrogen chloride.

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl\,(g)$$

5. Define chemical equilibrium.

Chemical equilibrium: When the rate of forward reaction becomes equal to the rate of backward reaction. This state is called chemical equilibrium.

III. Answer in brief.

1.Define exothermic reaction with an example.

A reaction which takes place with the evolution of heat is known as exothermic reaction. For an exothermic reaction, $H_p < H_R$. Hence, $\Delta H = -$ ve.

eg. (i) $C(s) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = -393.5$ kJ

eg: (ii) $S(s) + O_2(g) \longrightarrow SO_2(g)$; $\Delta H = -297$ kJ

2.The rate of various reactions are different. Give reason.

A chemical reaction involves the breaking and making of bonds. A strong bond requires more energy to break, than a weak bond. A reaction which involves breaking of strong bonds, will be slow or impossible to occur at room temperature. On the other hand, a reaction which involves breaking of only weak bonds will be fast under similar conditions.

3.What is an irreversible reaction. Give example.

A reaction in which the products formed do not recombine to produce the original reactants is called an **irreversible reaction**.

Reactions involving liberation of a gas or those in which a precipitate is formed are generally **irreversible reactions**.

eg.: 1. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$

2. $AgNO_3 + KCl \longrightarrow AgCl \downarrow + KNO_3$

4.How does the temperature influence the rate of the reaction?

Aim: To find out how temperature affects reaction speed.

1. Take 50 ml of hydrochloric acid in a boiling tube. Record the temperature with a thermometer.(27°C) Then add a piece of clean magnesium ribbon. Use the thermometer to gently stir the acid as the magnesium reacts. Record the time for the magnesium to dissolve in the acid.
2. Repeat the experiment, using the same volume of acid and the same length of magnesium ribbon, but first heat the acid to 40°C before adding the magnesium ribbon to the acid.

Fig. Temperature and Reaction speed

3. Compare the time taken for two reactions. Which reaction is faster? How does temperature affect the speed of reaction?

Results:

Experiment temperature	27°C	40°C
Time for magnesium to dissolve/seconds	50	20

The results show that reaction proceeds faster when the acid is heated to 40°C. This shows that the speed of reaction is greater at higher temperature.

In general, the speed of reaction doubles for every 10°C rise in temperature. So a reaction is approximately twice as fast at 60°C than at 50°C. A reaction at 70°C is $2 \times 2 = 4$ times faster than at 50°C.

5.What are the conditions for reversible reaction?

The following conditions are important for a reversible reactions:

1. The reaction should be done in a closed vessel.
2. None of the products should be removed from the vessel.
3. Temperature and pressure should be kept constant.

IV. Answer in detail.

1.The speed of the reaction depends on the temperature. Explain with an experiment.

Some chemical reactions required heating. How does temperature affect speed of a reaction? The following experiment shows how to find this out. Two reactions are carried out at different temperatures. All other conditions are kept the same.

Experiment: The reaction between magnesium ribbon and hydrochloric acid.

Aim: To find out how temperature affects reaction speed.

1. Take 50 ml of hydrochloric acid in a boiling tube. Record the temperature with a thermometer.(27°C) Then add a piece of clean magnesium ribbon. Use the thermometer to gently stir the acid as the magnesium reacts. Record the time for the magnesium to dissolve in the acid.
2. Repeat the experiment, using the same volume of acid and the same length of magnesium ribbon, but first heat the acid to 40°C before adding the magnesium ribbon to the acid.

Fig. Temperature and Reaction speed

3.Compare the time taken for two reactions. Which reaction is faster? How does temperature affect the speed of reaction?

Results:

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The results show that reaction proceeds faster when the acid is heated to 40°C. This shows that the speed of reaction is greater at higher temperature.

In general, the speed of reaction doubles for every 10°C rise in temperature. So a reaction is approximately twice as fast at 60°C than at 50°C. A reaction at 70°C is $2 \times 2 = 4$ times faster than at 50°C.

2.Describe the factors influencing the rate of a reaction.

Rate of a reaction depends on the following factors.

(a). Concentration of the reactants: Rate of a reaction increases with increase in concentration of the reactants. However, the rate constant does not change by changing the concentration of the reactant.

(b). Temperature: Any reaction becomes faster at a higher temperature. Increasing the temperature increases the rate of a reaction. Generally the rate of most chemical reactions increases by a factor of 2 or 3 times for a 10°C rise in temperature. Many biologically important reactions have greater temperature dependence. Their rates increase by a factor larger than 2 or 3 times when the temperature increases by 10°C.

(c). Presence of a Catalyst: A Catalyst is a substance that alters the rate of a chemical reaction but it is not used up in the course of a reaction. A positive catalyst increases the rate of the reaction whereas a negative catalyst decreases the rate.

(d). Nature of reactants: If the reactants are homogeneous, the reaction is faster. If they are heterogeneous the reaction is slower. *Eg:* Reaction between dilute hydrochloric acid and sodium carbonate solution is **faster** than the reaction between dilute hydrochloric acid and solid sodium carbonate.

(e). Nature of the solvent: It may be noted that many reactions are carried out in solutions. For these reactions, changing the solvent will generally change the rate of the reaction.

(f). Exposure to radiations: In some cases the rate of a chemical reaction is considerably increased by the use of certain radiations. The photons of these radiations having frequencies (υ) possess sufficient energies ($E = h\upsilon$) to break certain bonds in reactants. *For example* reactions of hydrogen and chlorine takes place very slowly in the absence of light. However in the presence of light, the reaction takes place very rapidly.

(g). Surface area: The larger, the surface area of the reactants, the faster is the rate of the reaction.

3.Write a note on classification of reaction.

On the basis of their rates (speed) chemical reactions can be broadly classified into three different types:

(a) Instantaneous (or) fast reactions: These reaction occur at once, for example, ionic reactions such as acid-base neutralization reaction and precipitation reaction

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

$$AgNO_3 + KCl \longrightarrow AgCl\downarrow + KNO_3$$

Reactions of this type proceed so quickly and their rates cannot be determined by common methods.

(b) Extremely slow reactions: Some reactions proceed at extremely slow rates. *For example*, rusting of iron, combination of hydrogen and oxygen at room temperature *etc.,* are few such reactions, which takes months, or even years before any observable change occur.

(c) Reactions with moderate speeds: In between the two extremes discussed above, there are a large number of reactions, which proceeds at moderate rates. A few examples of this type are given below.

(i) Decomposition of hydrogen peroxide

$$2H_2O_{2\,(liq)} \longrightarrow 2H_2O_{(liq)} + O_2(g)$$

(ii) Reaction of acetic acid with ethyl alcohol.

$$\underset{\text{Acetic acid}}{CH_3COOH_{(liq)}} + \underset{\text{Ethyl alcohol}}{C_2H_5OH_{(liq)}} \longrightarrow \underset{\text{Ethyl acetate}}{CH_3COOC_2H_{5\,(liq)}} + H_2O_{(liq)}$$

(iii) Acid catalysed hydrolysis of cane sugar in aqueous solution.

$$\underset{\text{Cane sugar}}{C_{12}H_{22}O_{11}} + H_2O \longrightarrow \underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Fructose}}{C_6H_{12}O_6}$$

4.Explain the exothermic and endothermic reaction with an experiment each.

Experiment: Exothermic and endothermic changes

Aim: To find out the heat change when salts dissolve in water.

Step 1 **Step 2** **Step 3**

Fig.

Step 1: Half-fill a boiling tube with tap water. Measure the temperature with a thermometer.

Step 2: Add a quantity of sodium carbonate. Shake the tube to dissolve the solid. Then measure the temperature of the solution. Has the temperature gone up or down?

Step 3: Repeat steps 1 and 2, using solid ammonium chloride in place of sodium carbonate.

With sodium carbonate, there is an increase in temperature. The solution gives heat to the atmosphere. So it must have lost heat energy. Hence the solution has less energy than the solid and water at the start. This is an example of an ***exothermic change***.

With ammonium chloride, there is a decrease in temperature. The solution becomes cold. It takes in heat from the atmosphere, by doing this, it gains heat. Hence the solution has more energy than the solid and water at the start. This is an example of an ***endothermic change***.

Chemical reactions can be compared to dissolving salts in water. Some reactions are exothermic. They become hot and lose heat to the surroundings. Some reactions are endothermic. They become cold and gain heat from the surroundings.

Exothermic reaction: A reaction which takes place with the evolution of heat is known as exothermic reaction. For an exothermic reaction, $H_p < H_R$. Hence, $\Delta H = -$ ve.

eg. (i) $C(s) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = -393.5$ kJ

eg: (ii) $S(s) + O_2(g) \longrightarrow SO_2(g)$; $\Delta H = -297$ kJ

Endothermic reaction: A reaction which takes place with the absorption of heat is called endothermic reaction. For an endothermic reaction, $H_p > H_R$. Hence $\Delta H = +$ve.

eg. (i) $N_2(g) + O_2(g) \longrightarrow 2NO_2(g)$; $\Delta H = +180.8$ kJ

eg: (ii) $C(s) + 2S(s) \longrightarrow CS_2(g)$; $\Delta H = +117$ kJ

Note: H_R = Enthalpy of reactant

H_p = Enthalpy of product

$\Delta H = H_p - H_R$

ΔH = change in enthalpy (heat content)

5.In a reaction A + B $\longrightarrow$ C the concentration of C increases from 0.8 mol lit^{-1} to 3.2 mol lit^{-1} in 40 seconds. What is the rate of reaction?

Solution:

Initial concentration = 0.8 mol litre^{-1}

Final concentration = 3.2 mol litre^{-1}

$\therefore$ change in molar concentration = dx

$= 3.2 - 0.8 = 2.4$ mol litre^{-1}

Time taken = t = 40 s.

$$\therefore \text{rate} = \frac{dx}{dt} = \frac{2.4}{40} = 0.06 \text{ mol litre}^{-1} \text{ s}^{-1}$$

Rate = 0.06 mol lit^{-1} s^{-1}

Textbook problems with solutions (Page 68 & 69).

Problem 1:

In a reaction $A + B \longrightarrow C$ the concentration of C increases from 0.4 mol. lit^{-1} to 0.8 mol. lit^{-1} in 20 seconds. What is the rate of the reaction?

Solution:

$$\text{Rate of the reaction} = \frac{\text{Change in concentration of product}}{\text{Time}}$$

$$= \frac{d[C]}{dt} = \frac{[0.8 - 0.4]\ \text{mol. lit}^{-1}}{20\ \text{seconds}}$$

$$= 0.4\ \text{mol. lit}^{-1}/20\ \text{seconds}$$

$$\text{Rate of the reaction} = 0.02\ \text{mol. lit}^{-1}.\text{sec}^{-1}$$

Problem 2:

Calculate the active mass of *(a)* 8.5g ammonia gas in a vessel of $2dm^3$ capacity *(b)* 1.4g nitrogen in a vessel of $5dm^3$ capacity. (***Hint:*** *dm^3 = lit*)

Solution:

$$\text{Active mass} = \frac{\text{No. of moles}}{\text{Volume of the containing vessel}}$$

(a) Molecular mass of NH_3 $= 14 + 3 = 17$

No. of moles of NH_3 $= \frac{8.5}{17} =$ 0.5 moles (using Hint 2)

$\therefore$ Active mass of NH_3 $= \frac{0.5}{2} = 0.25\ \text{mol dm}^{-3}$

(b) Molecular mass of N_2 $= 14 \times 2 = 28$

No. of moles of N_2 $= \frac{1.4}{28} =$ 0.05 moles

$\therefore$ Active mass of N_2 $= \frac{0.05}{5} = 0.01\ \text{mol dm}^{-3}$

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1. For the hypothetical reaction. $A \longrightarrow C$, the reaction rate 'r' in terms of the rate of change of concentration is given by
   (a) r = -d [A] / dt (b) r = ½ d [A} / dt (c) r = 1/3 d [A] /dt (d) r = - d[C] / dt
2. For the reaction, $L + M \longrightarrow X + Y$, the rate is given by
   (a) – d[L] / dt (b) + d[L] / dt (c) d [m] / dt (d) – d[X] / dt
3. For the reaction: $A + B \longrightarrow C$, the rate of reaction at a given instant of time can be represented by-
   (a) + d[A] /dt = - d[B] / dt = + d[C]/dt (b) - d[A] /dt = + d[B] / dt = - d[C]/dt
   (c) - d[A] /dt = - d[B] / dt = + d[C]/dt (d) + d[A] /dt = + d[B] / dt = + d[C]/dt
4. The rate of reaction does not depends upon:
   (a) Temperature (b) initial concentration (c) pressure (d) none
5. The unit of the rate of the reaction is
   (a) mol L^{-1} (b) mol L^{-1} S^{-1} (c) mol LS^{-1} (d) L mol^{-1} S^{-1}

6. For a reversible reaction at equilibrium,
   (a) there is no change in volume
   (b) the reaction is stopped completely
   (c) the rate of forward reaction is equal to the rate of reverse reaction
   (d) the forward reaction is faster than the reverse reaction.
7. The rate of chemical reaction generally,
   (a) increase with increase in temperature (b) increase with decrease in temperature
   (c) decrease with increase in temperature (d) does not change with temperature
8. In which reaction the equilibrium is established when ____
   (a) Concentrations of reactants and products are same
   (b) no generation of heat in the reaction
   (c) rates of forward and backward reactions are same.
   (d) the reaction inn forward and backward direction is stopped.
9. The reaction A $\rightleftharpoons$ B will be in equilibrium if______
   (a) A will change completely into B
   (b) 50% of A will change into B
   (c) 50% of B will change into A (d) rate of conversion of A into B and of B into A will be same.
10. When the system is in equilibrium, then______
   (a) mass of products is equal to the mass of reactants
   (b) the number of molecules of reactants and products are same.
   (c) the ratio of velocity of forward and backward reactions are in the ratio 1:1
   (d) none of these.
11. According to the law of mass action, the velocity of a reactions is proportional to the
   (a) volume of the vessel (b) equilibrium constant
   (c) nature of reactants (d) on the product of molar concentration of reactants.
12. The rate of a chemical reaction
   (a) increase as the reaction proceeds (b) decrease as the reaction proceeds
   (c) may increases as the reaction proceeds (d) remains constant as the reaction proceeds.
13. A mathematical way of representing the rate of a chemical reaction is
   (a) $\pm d_x \times d_t$ (b) $d_x + d_t$ (c) $d_x - d_t$ (d) d_x / d_t
14. The unit for stating the rate of a reaction is:
   (a) mol dm^{-3} (b) fast or slow (c) litres
   (d) mole per litre per minutae
15. Chemical reactions are in a state of dynamic equilibrium
   (a) the rate of the forward reaction equals that of the reverse reaction
   (b) the concentration of reactants and products are equal .
   (c) the reaction involves no enthalpy change
   (d) all of 1, 2, and 3
16. When a reversible chemical reaction is at equilibrium
   (a) the concentration of reactants and products remains equal
   (b) the forward reaction is unable to continue
   (c) the concentration of reactants and products remain constant.
   (d) the forward and reverse reaction process at different rates
17. For a reaction in equilibrium
   (a) there is no change

   (b) the reaction has stopped
   (c) the rate of forward reaction is equal to the rate of backward reaction
   (d) the forward reaction is faster than reverse reaction.
18. Which of the following factors influences the rate of chemical reaction
   (a) concentration (b) temperature (c) pressure
   (d) all of 1, 2, 3
19. The law / principle represented by the following equation: K = [product] / [reactant]
   (a) equilibrium law (b) rate law (c) law of mass action
   (d) Le Chatelier's principle
20. Precipitation Reaction is an example for….
   (a) Fast reaction (b) slow reaction
   (c) Reactions with moderate speeds (d) none of the above
21. Which one of the following reaction is extremely slow?
   (a) Rusting of iron (b) neutralization of acid and base
   (c) Decomposition of hydrogen peroxide (d) formation of hydrogen chloride
22. The reaction takes place only when the reacting substance after mixing is exposed to light
   (a) Extremely slow reaction (b) fast reaction
   (c) Photochemical reaction (d) neutralization reaction
23. In Photosynthesis carbondioxide and water to give
   (a) Acid (b) carbohydrates (c) proteins (d) None of the above
24. Reaction of acidic acid with ethyl alcohol to form
   (a) Ethyl acetate (b) glucose (c) methyl acetate (d) ether

25. The reaction between sodium carbonate and water is an example for
(a) Exothermic reaction (b) endothermic reaction (c) fast reaction (d) none
26. The value of ΔH for exothermic reaction is…
(a) Negative (b) positive (c) zero (d) none
27. $H_P > H_R$ is …………….
(a) an exothermic reaction (b) an endothermic reaction
(c) an slow reaction (d) an fast reaction
28. The unit for change in molar concentration is..
(a) mol dm^{-3} s^{-1} (b) mol dm^{-3} (c) mol dm^{+3} (d) mol dm^{3} s^{1}
29.If the chemical reaction takes place the molar concentration of products will be ……
(a) increases (b) decreases (c) no change (d) first increases then decreases
30.The fast reaction occur when the breaking of ……..
(a) Strong bonds (b) weak bonds (c) ironic bonds (d) covalent bonds
31. The reactants A and B gives the product C at the time 10 seconds the concentration of C increases from 0.2 mole lit^{-1} to 0.4 mol lit^{-1} the rate of reaction is ……
(a) 0.2 mol lit^{-1} s^{-1} (b) 0.4 mol lit^{-1} s^{-1} (c) 0.12 mol lit^{-1} s^{-1} (d) 0.02 mol lit^{-1} s^{-1}
32. The unit for specific reaction rate in $A \rightarrow B$ is……
(a) mol dm^{-3} (b) sec^{-1} (c) mol. Lit^{-1} sec^{-1} (d) mol. sec^{-1}
33. 17 g ammonia gas in the vessel of $4dm^3$ capacity so the active mass is
(a) 0.50 (b) 0.1 (c) 0.25 (d) 0.75
34. The concentration increases the rate of reaction is
(a) Decreases (b) increases (c) first increases and decreases (d) no change
35. The rate of reaction increase by the factor of 2 or 3 times for rise in ………
(a) 10^oC (b) 20^oC (c) 100^oC (d) 1^oC
36. In the case of reversible reaction
(a) temperature should be constant (b) pressure should be constant
(c) the reaction should be done in closed vessel (d) all of the above
37. Which one of the following is an example for irreversible reaction?
(a) Formation of ammonia (b) formation of sulphur trioxide
(c) formation of silver chloride (d) formation of phosphorus trichloride
38. The reaction is approximately……….. faster at 70^oC than at 50^oC.
(a) 2 time (b) 3 times (c) 4 times (d) 8 times
39. The time required (in seconds) for magnesium to dissolve in HCl at 27^oC is……….
(a) 20 (b) 30 (c) 40 (d) 50
40. The time required (in seconds) for magnesium to dissolve in HCl at 40^oC is………
(a) 20 (b) 30 (c) 40 (d) 50
41. A reaction in which the products formed react to give back the original substance is called a ………. reaction.
(a) irreversible (b) reversible (c) chemical (d) All of these
42. A reversible reaction can be represented by…………
(a) single head arrrow ($\rightarrow$) (b) single head upward arrow($\uparrow$)
(c) double headed arrow ($\rightleftarrows$ (d) None of these
43. In a reversible reaction, which reaction takes place?
(a) Forward reaction only (b) Backward reaction only
(c) Simultaneous conversion of reactants into products and vice-versa
(d) None of these
44. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$. The reaction is………
(a) reversible (b) irreversible (c) Both (*a*) and (*b*) (d) None
45. Formation of AgCl from $AgNO_3$ and KCl is a ……….
(a) irreversible reaction (b) precipitation reaction
(c) Both (a) and (b) (d) None of these
46. $A + B \rightleftarrows C + D$. In the initial stages, the rate of …….. is high
(a) forward reaction (b) backward reaction (c) reaction (d) Both (*a*) and (*b*)
47. When the rate of forward reaction becomes equal to the rate of backward reaction, ……….. is reached.
(a) slow reaction (b) Chemical reaction
(c) non-equilibrium (d) chemical equilibrium
48. At equilibrium the concentrations of the reactants and the products…………
(a) changes with time (b) do not change with time
(c) are directly proportional to time (d) None of these
49. At equilibrium……….
(a) $R_f = R_b$ (b) $k_f = k_b$ (c) Both (a) and (b) (d) None of these

50. Hydrogen combines with chlorine in the presence of light rapidly due to …………
(a) weak bonds present in H_2 and Cl_2
(b) sunlight acts as a catalyst
(c) the photons of radiations possess sufficient energy to break bonds

(d) None of these
51. $H_2 + Cl_2 \longrightarrow 2HCl$. This reaction takes place in the presence of light ………
(a) slowly (b) moderately (c) to completion (d) very rapidly
52. In the absence of sunlight, the reaction between hydrogen and chlorine takes place…..
(a) fastly (b) very fastly (c) moderately (d) very slowly
53. The rate is observed fastly when…………
(a) smaller surface area of the reactants are present
(b) reactants are heterogenous
(c) reactants are solids
(d) surface area of the reactants are larger
54. In general, changing the solvent changes………
(a) rate of the reaction (b) rate constant (c) molecularity (d) None of these
55. Many reactions are carried out in ………… form.
(a) solid (b) solution (c) gases (d) None of these
56. The reaction between solid sodium carbonate and hydrochloric acid is found to be…..
(a) faster (b) slower (c) never complete (d) very much faster
57. $A \longrightarrow B$. What is the rate equation for this reaction?
(a) Rate = $k\,[A]^2$ (b) Rate = $k\,[A]$ (c) Rate = $k\,[A]\,[B]$ (d) Rate = $k\,[A]^2[B]$
58. Under what conditions rate of the reaction is equal to rate constant?
(a) When molar concentrations of the reactants are unity
(b) When molar concentrations of the products are taken as unity
(c) When both concentrations of the reactants and products are taken as zero
(d) All of these
59. $A + B \longrightarrow C + D$. In this reaction [A] and [B] are one mol/dm^3, then…………
(a) Rate = order (b) Rate = Rate constant (c) Rate = molecularity (d) None of these
60. Calculate the number of moles of ammonia gas contaning 85 g in a container?
(a)0.5 mole (b)0.05 mole (c) 50 moles (d) 5 moles
61. When 5 moles of ammonia gas is present in a 2 dm^3 vessel, its active mass is ……
(a) 25 mol dm^{-3} (b) 0.25 mol dm^{-3} (c) 2.5 mol dm^{-3} (d) 5mol dm^{-3}
62. The number of moles of nitrogen gas having 14 g is…………
(a) 0.05 moles (b) 0.5 moles (c) 5 moles (d) 0.25 moles
63. The ΔH value for the formation of CO_2 is……….
(a) -297 kJ (b) -393.5 J (c) -393.kJ (d) None
64. The enthalpy change for the formation of SO_2 is ……….
(a) +297 kJ (b) -297 J (c) -297 kJ (d) All of these
65. $N_2 + 2O_2 \longrightarrow 2NO_2$. This reaction is …………
(a) endothermic reaction (b) an exothermic reaction
(c) Precipitation (d) adsorption
66. The formation of carbondisulphide (CS_2) is……….. reaction.
(a) an adsorption (b) an exotgermic reaction (c) an endothermic (d) precipitation
67. The enthalpy change for the formation of NO_2 is ………
(a) +117 kJ (b) +180.8 J (c) +180.8 kJ (d) None
68. The ΔH value for the formation of carbondisulphide is………..
(a) +180.8 kJ (b) -297 kJ (c) +117 kJ (d) None
69. Rate is given by…………..
(a) $\frac{\text{Time}}{\text{Concentration}}$ (b) $\frac{\text{Change in molar concentration}}{\text{Time}}$
(c) timx concentration (d) concentration – time
70.The change in molar concentration of one of the reactants or products in unit time is called………..
(a) order of a reaction (b) molecularity of a reaction
(c) rate of chemical reaction (d) All of these
71. The increase in molar concentation of the product is given………sign in their rates.
(a) negative (b) positive (c) Both+ve and –ve (d) no
72. A chemical reaction involves the breaking and making of …………..
(a) bond (b)Water molecule (c) reactants (d) products
73. A strong bond requires ……….. to break, than a weak bond.
(a) energy (b) less energy (c) least energy (d) more energy
74. At room temperature, strong bonds are broken………………
(a) faster rate (b) slower rate (c) very much faster rate (d) None of these
75. At room temperature, weak bonds are broken…………..
(a) fastly (b) slowly (c) Both (a) and (b) (d) All of these
76. 0.005 moles of nitrogen is kept in 5 dm^{3}vessel. Calculate its active mass?
(a) 0.01 mol dm^{-3} (b) 0.5 mol dm^{-3} (c) 0.001 mol dm^{-3} (d) 0.1 mol dm^{-3}
77. When changing the concentration of the reaction,……….. does not change.
(a) rate (b) rate constant(c) order (d) None
78. The rate of the reaction is increased by………..
(a) positive catalyst (b) negative catalyst (c) auto catalyst (d) promoter

79. Negative catalyst……….. the rate of chemical reaction
(a) increases (b) decreases (c)does not affect (d) All of these

80. The reaction is faster, only when the reactants are………….
(a) heterogeneous (b) homogeneous (c) colloidal state (d) solids.

81. Decomposition of PCl_5 is………. Reaction
(a) irreversible (b) non reversible (c) reversible (d) chemical

82. Formation of nitric oxide is ………. Reaction
(a) irreversible (b) non reversible (c) chain (d) reversible

83. The reaction proceeding from left to right is called a ……….. reaction
(a) backward (b) forward (c) parallel (d) fast

84. A backward or reverse reaction is one in which the reaction proceeds from………..
(a) right to left (b) left to right (c) both sides (d) None of these

85. The reaction in which the products formed do not combine to produce the original reactants is called ………. Raction.
(a) reversible (b) irreversible (c) fast (d) slow

86. Irreversible reactions generally involves…………..
(a) liberation of gas (b) formation or precipitate
(c) Both (a) and (b) (d) None of these

87. Reactions in between the fast and slow reactions takes place at
(a) very fast (b) very slow (c) extremely slow (d) moderate speeds

88. Combination of hydrogen and oxygen at roomtemperature is an example of………
(a) fast reaction (b) moderate reaction
(c) slow reaction (d) photochemical reactions

89.The reaction between $H_{2(g)} + Cl_{2(g)} \xrightarrow{\text{light}} 2HCl_{(g)}$ is a ………reaction
(a) redox (b) ionic (c) photochemical (d) None

90. When wood and paper are burnt…………… is liberated.
(a) heat (b) cooling (c) Both (a) and (b) (d) None

Answers:

1. (a) 2. (a) 3. (c) 4. (c) 5. (b) 6. (c) 7. (a) 8. (c) 9. (d) 10. (c) 11. (d) 12. (a) 13. (d) 14. (d) 15. (a) 16. (a) 17. (c) 18. (d) 19. (c) 20. (a) 21. (a) 22. (c) 23. (b) 24. (a) 25. (a) 26. (a) 27. (b) 28. (a) 29. (a) 30. (a) 31. (d) 32. (c) 33. (c) 34. (b) 35. (a) 36. (c) 37. (c) 38. (c) 39. (d) 40. (a) 41. (b) 42. (c) 43. (c) 44. (b) 45. (c) 46. (a) 47. (d) 48. (b) 49. (a) 50. (c) 51. (d) 52. (d) 53. (d) 54. (a) 55. (b) 56. (b) 57. (b) 58. (a) 59. (b) 60. (d) 61. (c) 62. (b) 63. (c) 64. (c) 65. (a) 66. (c) 67. (c) 68. (c) 69. (b) 70. (c) 71. (b) 72. (a) 73. (d) 74. (b) 75. (a) 76. (c) 77. (b) 78.(a) 79. (b) 80. (b) 81. (c) 82. (d) 83. (b) 84. (a) 85. (b) 86. (c) 87. (d) 88. (c) 89. (c) 90. (a)

II. ANSWER IN ONE OR TWO SENTENCES:

1.What is thermo chemistry?

Thermo chemistry is a branch of chemistry which deals with the study of heat change during various physical and chemical transformations.

2.Define endothermic reactions.

Endothermic reaction: A reaction which takes place with the absorption of heat is called endothermic reaction. For an endothermic reaction, $H_p > H_R$. Hence $\Delta H = +ve$.

eg. (i) $N_2(g) + O_2(g) \longrightarrow 2NO_2(g)$; $\Delta H = + 180.8$ kJ

eg: (ii) $C(s) + 2S(s) \longrightarrow CS_2(g)$; $\Delta H = + 117$ kJ

Note: H_R = Enthalpy of reactant
H_p = Enthalpy of product
$\Delta H = H_p - H_R$
ΔH = change in enthalpy (heat content)

3.What do you mean by active mass?

'Active mass' means effective molar concentration and in a dilute solution, it can be considered to be equal to molar concentration expressed in mol/dm^3 or partial pressure expressed in (N/m^2) in the case of gases.

4. What are photo chemical reactions? Give Example.

Those reactions which are taking place in the presence of sun light are called photo chemical reactions**.**

Examples: (i) The reaction between hydrogen and chlorine in the presence of sunlight to give hydrogen chloride. (ii) Photosynthesis reaction.

5. What are reactions with moderate speeds? Give Example.

The reactions whose rates are in between fast reaction and slow reaction are said reactions with moderate speeds.

Example: (i) Decomposition of hydrogen peroxide.

(ii) Reaction of acetice acid with ethyl alcohol forming ethyl acetate and water.

(iii) Acid catalysed hydrolysis of cane sugar in aqueous solution.

6. What are slow reactions? Give Example.

Reactions which proceed at extremely slow rates are called slow reaction.

Example: Rusting of iron, combination of hydrogen and oxygen at room temperature are a few reactions which takes months or even years before any observable change occur.

7. Give examples for fast reactions?

Ionic reactions such as acid-base neutralisation and precipitation reaction proceed so quickly and their rates cannot be determined by common methods.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

$$AgNO_3 + KCl \longrightarrow AgCl\downarrow + KNO_3$$

8. How do you classify chemical reactions?

Chemical reactions can be broadly classified into three types on the basis of their rates. They are *(i)* instantaneous or fast reactions *(ii)* extremely slow reactions *(iii)* reactions with moderate speeds.

9. A + B $\longrightarrow$ C + D. Write rate equation for this chemical reaction.

According to law of mass action, rate for this reaction is

Rate $\propto$ [A] [B] where [A] [B] represents molar concentrations of A and B.

Rate = k [A] [B] where 'k' is a constant called rate constant or velocity constant or specific reaction rate.

10. What is meant by active mass? Give its units.

Active mass means effective molar concentration. For solutions it is expressed in mol/dm^3 or partial pressure (N /m^2) in the case of gases.

$$\text{Active mass} = \frac{\text{No of moles}}{\text{Volume of the containing vessel}}$$

11. What is meant by Enthalpy (or) ΔH?

ΔH means change in enthalpy or heat content. It is obtained from H_P and H_R.

$$\Delta H = H_P - H_R$$

where H_R represents enthalpy of reactant and H_P represents enthalpy of product.

12. Derive the unit of rate.

$$\text{Rate} = \frac{dx}{dt} = \frac{\text{mol / dm}^3}{\text{second}}$$

$\therefore$ the unit is mol / dm^3 / s or mol dm^{-3} s$_{-1}$.

13. How is rate represented? What is its unit?

Rate is represented as,

$$\text{Rate} = \frac{dx}{dt} = \frac{\text{Change in molar concentration}}{\text{Time}}$$

Its units is mol/dm^3 /s or mol dm^{-3} s^{-1}.

14. Give the significance of +ve and –ve sign given in the rate equation?

The +ve sign indicates that the molar concentration of product gradually increases with time and –ve sign indicates that the molar concentration of the reactant gradually decreases with time.

15. A + B $\longrightarrow$ C + D. write rate for this general reaction.

$$\text{Rate} = \frac{-d[A]}{dt} = \frac{-d[B]}{dt} = \frac{+d[C]}{dt} = \frac{+d[D]}{dt}$$

16. What is the influence of surface area on rate of the reaction?

The larger the surface area of the reactants, the faster is the rate of the reaction.

17. When is rate constant equal to rate of the reaction?

The rate constant is equal to rate of the reaction when the molar concentrations of the reactants are unity.

18. What is the effect of solvent on rates?

Many reactions are carried out in solutions. Changing the solvent will generally change the rate of the reaction.

19. Mention the factors influencing the rate of a reaction.

Rate of a reaction depends on the following factors:

(a) concentration of the reactants (b) temperature (c) presence of a catalyst (d) nature of reactants (e) nature of the solvent (f) exposure to radiations (g) surface area.

20. How is rate of the reaction affected by nature of the reactants?

If the reactants are homogeneous, the reaction is faster. If they are heterogeneous the reaction is slower.

21. How does the concentration influence the rate of the reaction?

Rate of a reaction increases with increase in concentration of the reactants. However the rate constant does not change by changing the concentration of the reactant.

22. What is the effect of catalyst on rate of the reaction?

A catalyst is a substance that alters the rate of a chemical reaction but it is not used up in the course of a reaction. A positive catalyst increases the rate of the reaction whereas a negative catalyst decreases the rate.

23. Give equations for two photochemical reactions.

(1) $H_{2(g)} + Cl_{2(g)} \xrightarrow{\text{light}} 2HCl_{(g)}$ (2) $6CO_2 + 6H_2O \xrightarrow{\text{light}} C_6H_{12}O_6 + 6O_2$

24. What is forward reaction? Give example.

In a reversible reaction, the reaction proceeding from left to right is called a forward reaction.

$C + O_2 \longrightarrow CO_2$

25. A + B $\rightleftarrows$ C + D. Represent the equilibrium diagrammatically.

DIAGRAM Page 4.15 Ide

26. What do you mean by reverse reaction? Give example.

The reaction proceeding from right to left is called reverse reaction or backward reaction.

Ex: $H_2 + I_2 \rightleftarrows 2HI$

III. ANSWER IN BRIEF:

1.What is reversible reactions? Give some examples.

A reaction in which the products formed react to give back the original substances is called a reversible reaction.

In a reversible reaction, the reactants are changed into products and simultaneously the products are changed into reactants. A reversible reaction is represented by as in,

$A + B \rightleftarrows C + D$

Examples:

1. $N_2 + 3H_2 \rightleftharpoons 2NH_3$
2. $2HI \rightleftharpoons H_2 + I_2$
3. $N_2 + O_2 \rightleftharpoons 2NO$
4. $2SO_2 + O_2 \rightleftharpoons 2SO_3$
5. $PCl_5 \rightleftharpoons PCl_3 + Cl_2$

In a reversible reaction, the reaction proceeding from left to right is called a forward reaction. The reaction proceeding from right to left is called a backward reaction or reverse reaction.

2. Write a short note on chemical equilibrium.

Consider a reversible reaction, $A + B \rightleftharpoons C + D$. When A and B are mixed in a closed vessel, they react to form C and D. Then, C and D start reacting back to form A and B. However, in the initial stages, the rate of forward reaction is high. Gradually, when the concentrations of A and B decrease, the rate of forward reaction decreases.

On the other hand, the rate of backward reaction gradually increases with an increase in amounts of C and D. After sometime, a state is reached when the rate of forward reaction becomes equal to the rate of backward reaction. This state is called chemical equilibrium.

Chemical equilibrium is a state in a reversible reaction when the rate of forward reaction becomes equal to the rate of backward reaction. The concentrations of the reactants and the products do not change with time in this state.

Fig. Chemical Equilibrium

ADDITIONAL PROBLEMS

1. In reaction A + B + C → D the concentration of D increases from 0.4 mol. lit^{-1} to 2 mol. lit^{-1} in 20 seconds. What is the rate of reaction?

Solution:

$$\text{Rate of reaction} = \frac{\text{Change in molar concentration}}{\text{Time}}$$

$$= \frac{d[c]}{dt} = \frac{20 - 0.4}{20}$$

$$\text{Rate} = 0.08 \text{ mol litre}^{-2} \text{ s}^{-1}$$

2. 0.5 lit. of solution containing 10g of sodium hydroxide, calculate the active mass of sodium hydroxide.

Solution:

$$\text{Active mass} = \frac{\text{No. of moles}}{\text{Volume of solution}}$$

$$\text{No. of moles} = \frac{\text{Weight}}{\text{Molecular weight}} = \frac{10}{40}$$

$$= 0.4 \text{ moles}$$

$$\text{Active mass} = \frac{0.4}{0.5} \text{ moles / lit.}$$

$$= 0.8 \text{ moles / lit.}$$

3. The concentration of reactant (A) changes from 0.06 to 0.02m in 40 minutes, calculate the rate of the reaction.

Solution:

$$\text{Rate of reaction} = \frac{\text{Change in concentration of reactant}}{\text{Time}}$$

$$= -\frac{d[A]}{dt}$$

$$= -\frac{[0.06 - 0.02]}{40} = 0.001$$

$$= 0.001 \text{ mol. lit}^{-1} \text{ sec}^{-1}$$

4. 200 ml of 0.5 moles of ammonia is present in a vessel. Calculate the active mass of ammonia.

Solution:

$$\text{Active mass} = \frac{\text{No. of moles}}{\text{Volume of vessel}}$$

$$= \frac{0.5}{0.2} = 2.5$$

$$= 2.5 \text{ mol. lit}^{-1}$$

5. Calculate the active mass of 88 g of carbon dioxide in 2 dm^{-3} vessel.

Solution:

$$\text{Active mass} = \frac{\text{Number of moles}}{\text{Volume of the vessel}}$$

$$\text{Molecular mass of } CO_2 = 12 + 32 = 44$$

$$\text{No. of moles of } CO_2 = \frac{88}{44} = 2 \text{ moles}$$

$$\text{Active mass of } CO_2 \quad \frac{2}{2} = 1$$

$$\text{Active mass of } CO_2 = 1 \text{ mol dm}^{-3}.$$

6. Calculate the active mass 36 gm of hydrogen in 4 dm^3 vessel.

Solution:

$$\text{Active mass} = \frac{\text{Number of moles}}{\text{Volume of the vessel}}$$

$$\text{Molecular mass of } H_2 = 2$$

$$\text{No. of moles of } H_2 = \frac{36}{2} = 16 \text{ moles}$$

$$\text{Active mass} = \frac{n}{v} = \frac{16}{4} = 4 \text{ mol dm}^{-3}$$

$$\text{Active mass of } H_2 = 4 \text{ mol dm}^{-3}.$$

7. For a general reaction A $\longrightarrow$ Products, the rate of the reaction is rate = k [A]. The initial rate of the reactant is 6 x 10^{-6} mol lit^{-1} s^{-1} at 298 K. The initial concentration of the reactant is 3 x10^{-3} m. Calculate the rate of constant of the reaction at 298 K.

$$\text{Rate} = k\,[A] \qquad \ldots\ldots 1$$

$$\text{Given A} = 3 \times 10^{-3} \text{ m}$$

$$\text{Rate} = 6 \times 10^{-6} \text{ mol}^{-1} \text{ lit}^{-1} \text{ s}^{-1}$$

Substituting the values of A and rate in equation 1, we get

$$6 \times 10^{-6}\ mol^{-1}\ lit^{-1}\ s^{-1} = k \times 3 \times 10^{-3}\ m$$

$$k = \frac{6 \times 10^{-6}\ mol^{-1}\ lit^{-1}\ s^{-1}}{3 \times 10^{-3}\ m}$$

$$= 2 \times 10^{-3}\ sec^{-1}$$

8. For a reaction X $\longrightarrow$ Products, the rate of the reaction is rate = k [X]. The rate constant of the reaction is 4 x 10^{-3} s^{-1} at 298 K. The initial concentration of the reactant is 0.08 mol lit^{-1}. Calculate the initial rate of the reaction at 298 K.

Solution:

$$\text{Rate} = k \times [A]$$
$$= 4 \times 10^{-3}\ s^{-1} \times 0.08\ mol\ lit^{-1}$$
$$= 0.32 \times 10^{-3}\ mol\ lit^{-1}\ s^{-1}$$

9. The decomposition of N_2O_5 is a first order reaction. When the concentration of N_2O_5 is 0.16 mol dm^{-3} the rate of reaction is found to be 0.056 mol dm^{-3} mm^{-1}. What is the rate of the reaction when the concentration is 0.100 mol dm^{-3}.

Solution:

Since it is a first order reaction, rate $= k\,[N_2O_5]$

$$\therefore k = \frac{\text{rate}}{[N_2O_5]} = \frac{0.056}{0.16} = 0.35\ min^{-1}$$

When concentration $= 0.100\ mol\ dm^{-3}$

rate = k x concentration = 0.35 x 0.100 $= 0.035\ mol\ dm^{-3}\ min^{-1}$

10. 5 moles per litre of $ZnCO_3$ are heated. After 9 seconds, the concentration

Solution:

We known that rate of reaction $= \dfrac{\text{Number of moles per litre of reactant consumed}}{\text{Time taken}}$

Given Time $= 9 - 0 = 9$ seconds

Number of moles per litre of $ZnCO_3$ consumed $= 5 - 2 = 3$

Substituting the values, we get rate of reaction $= \dfrac{3M}{9\ s} = 0.33\ ms^{-1}$.

11. 8 mol lit^{-} of $BaCO_3$ are heated. After 4 seconds the concentration of $BaCO_3$ left is 4 mol lit^{-}. Calculate the rate of reaction.

We known that rate of reaction $= \dfrac{\text{Number of moles per litre of reactant consumed}}{\text{Time taken}}$

Time = 4 seconds

Number of moles per litre of reactant consumed $= 8 - 4 = 4$

Substituting the values, rate of reactions $= \dfrac{4}{4} = 1\ mole\ sec^{-1}$

12. If the concentration of X and Y are expressed in terms of mole dm-3 and time in minutes, calculate the units for the rate constant for the following reaction X + Y $\longrightarrow$ XY.

Solution:

For the reaction, Rate $= k [X] [Y]$ ……1

Substituting the units of rate, X and Y in relation (1), we get

$$\frac{\text{mol dm}^{-3}}{\text{min}} = k\ (\text{mol dm}^{-3})\ (\text{mol dm}^{-3})$$

or $k = (\text{mol dm}^{-3})^{-1}\ (\text{min})^{-1} = \text{mol}^{-}\ \text{dm}^{3}\ \text{min}^{-1}$

13. The reaction $N_2O_5 \longrightarrow 2NO_2 + 1/2O_2$ is of first order in N_2O_5. Its rate constant is 6.2×10^{-6} s^{-1}. If in the beginning, $[N_2O_5]$ is 15 mol lit^{-1}, calculate the rate of reaction in the beginning.

Solution:

$k = 6.2 \times 10^{-6}\ s^{-1}$; $[N_2O_5] = 15$ mol lit^{-1}

Rate of reaction $= k [N_2O_5] = 6.2 \times 10^{-6}\ s^{-1} \times 15$ mol lit^{-1}

$= 93 \times 10^{-6}$ mol $lit^{-1}\ s^{-1} = 9.3 \times 10^{-5}$ mol $lit^{-1}\ s^{-1}$

14. The rate of a first order reaction, when the concentration of reactant is 10^{-1} mol, lit^{-1}, is 3×10^{-4} mol lit^{-1} sec^{-1}. What will be the rate of the reaction when the concentration of reactant is 10^{-2} mol lit^{-1}?

Solution:

For a first order reaction $A \longrightarrow$ Products; rate $= k[A]$

$$k = \frac{\text{rate}}{[A]} = \frac{3 \times 10^{-4}\ \text{mol lit}^{-1}\ \text{sec}^{-1}}{10^{-1}\ \text{mol lit}^{-1}} = 3 \times 10^{-3}\ \text{sec}^{-1}$$

$\therefore$ Rate or reaction when the concentration is 10^{-2} mol lit^{-1}

$= k [A] = 3 \times 10^{-3} \times 10^{-2}$ mol $lit^{-1}\ sec^{-1}$

$= 3 \times 10^{-5}$ mol $lit^{-1}\ sec^{-1}$

15. Calculate the active mass of oxygen if 6.4 g of oxygen is held in a 250 ml vessel.

Active mass of oxygen,

$$[O_2] = \frac{\text{No. of moles of oxygen}}{\text{volume in lit (dm}^3)}$$

$$\text{No. moles of oxygen} = \frac{\text{Weight of oxygen}}{\text{gram molecular weight}}$$

$$= \frac{6.4}{32} = 0.2 \text{ moles}$$

$$\text{Volume} = 250 \text{ ml} = \frac{250}{1000} = 0.25 \text{ lit} = 0.25 \text{ dm}^{-3}$$

$$\therefore [O_2] = \frac{0.2}{0.25} = 0.8 \text{ moles dm}^{-3}$$

16. Calculate the active mass of nitrogen if 56 g of nitrogen is held in a 200 ml container.

Active mass of nitrogen,

$$[N_2] = \frac{\text{No. of moles of nitrogen}}{\text{volume in lit } (dm^3)}$$

$$\text{No. moles of nitrogen} = \frac{\text{Weight of nitrogen}}{\text{gram molecular weight}}$$

$$= \frac{56}{28} = 2 \text{ moles}$$

$$\text{Volume} = 200 \text{ ml} = \frac{200}{1000} = 0.20 \text{ lit} = 0.20 \text{ dm}^{-3}$$

$$\therefore [N_2] = \frac{2}{0.20} = 10 \text{ moles dm}^{-3}$$

17. Calculate active mass of 96 g of oxygen contain in 2 lire flask.

Active mass of oxygen,

$$[O_2] = \frac{\text{No. of moles of oxygen}}{\text{volume in lit } (dm^3)}$$

$$\text{No. moles of oxygen} = \frac{\text{Weight of oxygen}}{\text{gram molecular weight}}$$

$$= \frac{96}{32} = 3 \text{ moles}$$

$$\text{Volume} = 2 \text{ lit} = 2 \text{ dm}^{-3}$$

$$\therefore [O_2] = \frac{3}{2} = 1.5 \text{ moles dm}^{-3}$$

18. Calculate the active mass of nitric oxide (NO) when 90 grams of it is held in a two litre flask.

Active mass of nitric oxide,

$$[NO] = \frac{\text{No. of moles of nitric oxide}}{\text{volume in lit } (dm^3)}$$

$$\text{No. moles of nitric oxide} = \frac{\text{Weight of nitric oxide}}{\text{gram molecular weight}}$$

$$= \frac{90}{30} = 3 \text{ moles}$$

$$\text{Volume} = 2 \text{ lit} = 2 \text{ dm}^{-3}$$

$$\therefore [NO] = \frac{3}{2} = 1.5 \text{ moles dm}^{-3}$$

19. What is the active mass of hydrogen when 20 g hydrogen gas is present in 1 dm^{-3} flask?

Active mass of hydrogen,

$$[H_2] = \frac{\text{No. of moles of Hydrogen}}{\text{volume in lit } (dm^3)}$$

$$\text{No. moles of Hydrogen} = \frac{\text{Weight of hydrogen}}{\text{gram molecular weight}}$$

$$= \frac{20}{2} = 10 \text{ moles}$$

$$\text{Volume} = 1 \text{ lit} = 1\ dm^{-3}$$

$$\therefore [H_2] = \frac{10}{1} = 10 \text{ moles } dm^{-3}$$

20. 85 g ammonia is present in 2 litre vessel. Calculate the molar concentration of ammonia.

Active mass of ammonia,

$$[NH_3] = \frac{\text{No. of moles of ammonia}}{\text{volume in lit } (dm^3)}$$

$$\text{No. moles of ammonia} = \frac{\text{Weight of ammonia}}{\text{gram molecular weight}}$$

$$= \frac{85}{17} = 5 \text{ moles}$$

$$\text{Volume} = 2 \text{ lit} = 2\ dm^{-3}$$

$$\therefore [NH_3] = \frac{5}{2} = 2.5 \text{ moles } dm^{-3}$$

21. Following reaction was carried out at 300K $2SO_{2(g)} + O_{2(g)} \longrightarrow 2SO_{3(g)}$ How is the rate of formation of SO_3 related to the rate of disappearance of O_2 is

Solution: Rate of reaction

$$= -\frac{1}{2}\frac{d[SO_2]}{dt} = -\frac{d[O_2]}{dt} = +\frac{1}{2}\frac{d[SO_2]}{dt}$$

Therefore rate of disappearance of O_2 is related to rate of formation of SO_3 as

$$-\frac{d[O_2]}{dt} = +\frac{1}{2}\frac{d[SO_3]}{dt}$$

22. The rate of change in concentration of C in the reaction $2A + B \longrightarrow 2C + 3D$ was reported 1.0M Sec^{-1}. Calculate the reactions rate as well as rate of change of concentration of A, B and D.

Solution: We have,

$$= -\frac{1}{2}\frac{d[A]}{dt} = -\frac{d[B]}{dt} = +\frac{1}{2}\frac{d[C]}{dt} = -\frac{1}{3}\frac{d[D]}{dt} = \text{rate of reaction}$$

$$\therefore\ -\frac{d[C]}{dt} = 1.0 \text{ mol lit}^{-1} \text{ Sec}^{-1}$$

$$\therefore\ -\frac{d[A]}{dt} = \frac{d[C]}{dt} = 1.0 \text{ mol lit}^{-1} \text{ Sec-1}$$

$$\therefore\ -\frac{d[B]}{dt} = \frac{1}{2}\frac{d[C]}{dt} = 0.5 \text{ mol lit}^{-1} \text{ Sec}^{-1}$$

$$\therefore\ -\frac{d[D]}{dt} = \frac{3}{2}\frac{d[C]}{dt} = \frac{3}{2} \times 1 = 1.5 \text{ mol lit}^{-1} \text{ Sec}^{-1}$$

also,

$$\therefore\ \text{Rate} = \frac{1}{2}\frac{d\,[C]}{dt}$$

$$\text{Rate} = \frac{1}{2} \times 1 = 0.5 \text{ mol lit}^{-1} \text{ sec}^{-1}$$

5- METALLURGY

INTRODUCTION:

The entire matter on this earth is made up of different elements. Elements are classified as metals, non-metals and metalloids.

Metals constitute about two third of the known elements. Metals are generally solids and are malleable and ductile. They possess characteristic lusture and are good conductors of heat and electricity. Most metals occur in the combined state. Only unreactive metals are lound in the free state or native form e.g., silver, gold and platinum.

Non-metals can be solids, liquids or gaseous. They are non-lustrous, brittle and poor conductors of heat and electricity. (But graphite is a good conductor). Metalloids show characteristics of metals and non-metals. Antimony, arsenic, etc. are metalloids.

A naturally occurring compound of a metal is called a mineral. **The mineral from which the metal can be extracted profitably is known as an ore**. The science of extracting the metals from their ores is called metallurgy.

The whole process of extraction of metal in the free state from its ore is called metallurgy. The common steps involved in the metallurgical process are the following.
(i) Crushing and pulvarisation of the ore
(ii) Concentration of the ore
(iii) Extraction of metal from the concentrated ore
(iv) Refining of the crude metal.

IMPORTANT TERMS & DEFINITIONS

1. **Metallic lustre**: The property of metals due to which they shine is called metallic lustre. The bright lustre of metals is due to their ability to reflect the incident light from their face and the surface acquires a shining appearance, which is known as metallic lustre.
2. **Metals are solids:** Most of the metals are generally solids at room temperature except mercury, which is liquid at room temperature.
3. **Metals are hard:** Due to strong forces of attraction between the metal atoms, they are generally hard except sodium and potassium which are soft metals .The hardness varies from metal to metal. The metals like magnesium (Mg), lead (Pb), aluminium (Al), Iron (Fe), copper (Cu), etc. cannot be cut with a knife, so these are not soft metals.
4. **Metals are malleable and ductile:** Metals can be beaten into sheets (malleability) and drawn into wires (ductility).
5. **Thermal conductivity:** Metals are good conductors of heat. For example iron, silver, gold, aluminium , etc. conduct heat.
6. **Electrical conductivity:** Metals are good conductors of electricity.
7. **Sonorous:** Most of the metals produce ringing sound when they are struck.
8. **Tensile strength:** Metals resist breaking when stretched. This is a measure of their tensile strength. Metals like tungsten have very high tensile strength.
9. **Density:** Metals generally have high densities, e.g., density of gold, mercury and iron is 19.3, 13.6 and 7.6 g/cc. respectively.

Minerals: The various compounds of metals, which occur in nature and are obtained by mining, are called minerals.
Ores: These are minerals from which metals can be conveniently and economically extracted.
Matrix or gangue: The unwanted impurities such as mud, stones, sand etc. which are present in the ore are called matrix or gangue.
Metallurgy: The process of extraction of pure metals from their ores is called metallurgy. The method of extraction of metal depends on the nature of the metal and the nature of its ore.
Refining of metals: The process of purification of impure metals by removing metallic and nonmetallic impurities is known as refining of metals.
Poling: The impure metal is melted and the molten metal is stirred with logs of green wood. The impurities are removed either as gases or they get oxidised forming scum over the molten metal
Liquation: This method is used for refining those metals, which have low melting point such as tin, lead etc., The impure metal is placed on the sloping hearth of a furnace and gently heated. The metal melts and drain away leaving behind the infusible materials on the hearth.

Distillation: This process is employed for purification of volatile metals like mercury, zinc and cadmium. The impure metal is heated in a retort and its vapours are separately condensed in a receiver. While the pure metal distills over, the non-volatile impurities are left behind in the retort.
Zone refining: Ultra pure metals and non-metals are obtained by zone refining process. It is also called fractional crystallisation method because this refining is based on the principle that when an impure metal is melted and allowed to solidify, the impurities move away from the solid region and prefer to be distributed in the molten region.
Electrolytic refining: This method is most widely used for refining impure metals. Metals such as copper, zinc, tin, nickel etc., are refined electrolytically.
Oxidation: This method is generally employed in the purification of metals, when the impurities get oxidised more readily than the metal itself.

SELF EVALUATION (T.B.PAGE. 88)

I. Choose the correct answer.

1. The gas evolved during the reaction of sodium with water is ....
   (a) H_2 (b) N_2
   (c) NaH (d) None of the above
2. Which one of the following metal occurs in native form?
   (a) Platinum (b) Iron
   (c) Aluminium (d) Zinc
3. The process of extraction of pure metals from their ores is known as ....
   (a) Electro refining (b) Metallurgy
   (c) Enrichment (d) Electrolysis
4. .... is the process used to eliminate the impurities in gaseous form or by forming scum over the molten metal.
   (a) Liquation (b) Poling
   (c) Distillation (d) Oxidation
5. Germanium is purified by .... method.
   (a) Zone refining (b) Distillation
   (c) Oxidation (d) Liquation
6. The volatile metals can be purified by .... method.
   (a) Zone refining (b) Distillation
   (c) Oxidation (d) Liquation
7. The process used for removing the gangue from the ore is known as ....
   (a) Metallurgy (b) Enrichment
   (c) Refining (d) Electrolysis
8. Name the metal, which do not react with steam.
   (a) Gold (b) Aluminium
   (c) Iron (d) Magnesium
9. .... is the gangue present in the haematite ore.
   (a) Ferrous Oxide (b) Copper sulphide
   (c) Calcium Sulphide (d) Calcium Silicate
10. Which among the following metal is enriched by electromagnetic separation?
   (a) Copper (b) Tin
   (c) Silver (d) Iron

Answers:

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

II. Answer the following in One or Two sentences. (T. B. Page 89)

1.What do you mean by metallic lustre?

The property of metals due to which they shine is called metallic lustre. The bright lustre of metals is due to their ability to reflect the incident light from their face and the surface acquires a shining appearance, which is known as metallic lustre. Iron, silver, magnesium, aluminium etc. appear as white whereas gold is yellow and copper is reddish brown in appearance.

2.Define matrix or gangue.

The unwanted impurities such as mud, stones, sand etc. which are present in the ore are called matrix or gangue.

3.What is liquation?

This method is used for refining those metals, which have low melting point such as tin, lead etc the impure metal is placed on the sloping hearth of a furnace and gently heated. The metal melts and drain away leaving behind the infusible materials on the hearth.

4.What is meant by hydraulic wash?

This method is suitable for enrichment of heavy oxide ore. In this the powdered ore is placed on a sloping surface and washed in a strong current of water. The heavier metallic ore particles settle down at the bottom and the lighter impurities are washed away.

5.$Na + Cl_2 \longrightarrow$. Complete and balance the reaction.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

$$Ca + Cl_2 \longrightarrow CaCl_2$$

6.Potassium reacts with water more readily than sodium - Why?

The reactivity of potassium with water is more than that of sodium. This is because, the reaction of potassium with water is so violent that the evolved hydrogen catches fire.

III. Answer in brief.(T.B. Page 89)

1.Define Calcination with an example.

Calcination is the process of conversion of ore into metal oxide (oxidation) by heating strongly in absence of excess of air at high temperature.

During calcination the volatile impurities are removed and the mass become porous. Carbonate and oxide ore are generally calcinated where they lose moisture, carbon dioxide and other volatile impurities.

$$\underset{\text{Malachite}}{CuCO_3 .Cu(OH)_2} \longrightarrow \underset{\text{Copper oxide}}{2CuO} + H_2O + CO_2\uparrow$$

The metal oxides are then reduced to the corresponding metals by using suitable reducing agent.

2.What do you mean by electromagnetic separation of ores?

This method is used for separating magnetic impurities from non-magnetic ore particles. *For example,* Tin stone (tin ore) in which tinstone is non-magnetic containing wolfromite as magnetic particle. The powdered ore is dropped over the moving belt passing over the electromagnetic roller. Wolfromite being paramagnetic is attracted by the magnet and forms a heap nearer to the roller, while tin stone fall away from the roller and forms another separate heap.

Fig. Magnetic Separation

3.Explain the action of oxygen on metals with two example.

Metals generally combine with oxygen to from oxides, which are basic in nature.

Example: 1 $2Zn + O_2 \longrightarrow 2\,ZnO$

Zinc oxide

*Example: 2*2 $Mg + O_2 \longrightarrow 2\,Mg\,O$

Magnesium oxide

4.Metals are malleable and ductile – Justify.

Metals can be beaten into sheets (malleability) and drawn into wires (ductility). Metals like Silver, Iron, Copper and Aluminium expand when striked with a hammer. Gold is the most malleable metal. Metals like gold, aluminium, copper silver, *etc.,* can be drawn into thin wires very easily and are, therefore, said to be ductile and this property of metals is called ductility.

5.Write a note on electrolytic refining?

This method is most widely used for refining impure metals. Metals such as copper, zinc, tin, nickel etc., are refined electrolytically.

The impure metal to be refined is made, as anode of an electrolytic cell while the cathode is a thin plate of the pure metal. On passing the electric current, the impure metal dissolves, go into solution while insoluble matter settles down at the bottom and is called the anode mud.

IV. Answer in detail. (T.B. Page 90)

1.Distinguish metals and non-metals based on their physical properties.

S. No.	Characteristics	Metals	Non-metal
1.	Physical state	They are solid at room temperature except Hg which is a liquid.	They are either solids or gases except bromine which is a liquid.
2.	Density	They usually have high density. Exception: Na and K are metals but their densities are less than that of water.	They usually have low density.
3.	Conductivity	They are good conductors of heat and electricity except bismuth.	They are bad or poor conductors of heat and electricity, except graphite, which is a good conductor.
4.	Metallic lustre	They are lustrous (shining).	They are non-lustrous (dull), except graphite and iodine.
5.	Malleability	They are malleable.	They are non-malleable and are brittle.
6.	Melting and boiling point	They have high melting and boiling points, except Na and K.	They have low melting and boiling point except graphite.
7.	Ductility	They are ductile.	They are non-ductile.
8.	Sonorous	They are sonorous.	Non-sonorous.
9.	Tensile strength	They have high tensile strength.	They have low tensile strength.
10.	Hardness	They are hard except Na and K, which are soft metals.	They are soft except diamond, which is hard.

2.Distinguish metals and non-metals based on their chemical properties.

S. No.	Metals	Non-metal
1.	Metals are electro positive. *Example:* Na^+, K^+, Ca^{2+}, *etc.*	Nonmetals are electro negative. *Example:* Cl^-,S^{2-},N^{3-}, *etc.*
2.	Oxides of metals are basic in nature, *i.e.,* the oxides of metals react with water to give bases or alkalis. $Na_2O + H_2O \longrightarrow 2\ NaOH$	Oxides of nonmetals are acidic in nature, i.e., the oxides of nonmetals react with water to give acids. $SO_3 + H_2O \longrightarrow \underset{\text{Sulphuric acid}}{H_2SO_4}$
3.	Metals dissolve in dilute acids to produce hydrogen gas. $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$	Nonmetals generally do not dissolve in dilute acids.

4.	Metals in general do not combine with hydrogen. Exceptions Some metals (Na, Ca , Li *etc.*) combine with hydrogen to form non-volatile unstable hydrides.	Nonmetals combine with hydrogen to form stable hydrides. *eg:* NH_3, H_2S, PH_3
5.	Metallic chlorides are generally not hydrolysed by water, or are only partially hydrolysed. $NaCl + H_2O \rightarrow$ No hydrolysis $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$	Chlorides of nonmetals are usually hydrolysed by water. $PCl_3 + 3H_2O \rightarrow 3HCl + H_3PO_3$ $SiCl_4 + 4H_2O \rightarrow Si(OH)_4 + 4HCl$
6.	Metals are reducing agents.	Nonmetals are oxidizing agents, except carbon, which is a reducing agent.

3.Explain the action of water on metals with suitable chemical equations.

Action with water on metals: Metals on reaction with water form metal oxide or metal hydroxide and liberate hydrogen gas.

(i) Sodium and potassium react vigorously with cold water forming their hydroxides and liberating H_2 gas.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

The reactivity of potassium with water is more than that of sodium. This is because, the reaction of potassium with water is so violent that the evolved hydrogen catches fire.

(ii) Magnesium on reaction with water form magnesium oxide and hydrogen.

$$Mg + H_2O \longrightarrow MgO + H_2$$

(iii) Heated iron reacts with steam to give ferroso ferric oxide and hydrogen.

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

(iv) Metals like Au, Ag and Cu do not react even with steam.

4.Describe the enrichment of ores by

***(a)* Froth flotation process.**

***(b)* Electrolysis.**

(a) Froth floatation process: This process is generally used for the concentration of sulphide ores. In this method the ore is taken in an iron tank along with a mixture of water and pine oil. The mixture is vigorously agitated by blowing compressed air. The oil forms a froth with air. The metallic ore particles get stick to the forth and rise to the surface as scum.

The impurities are wetted by water and will settle down at the bottom .The froth containing the metallic ore particles is skimmed off and dried.

Fig. Froth Floatation Process

(b) Chemical Method: This method is generally used in the case where the ore is to be in a very pure form. *e.g.* Aluminium extraction.

Bauxite $Al_2O_3.2H_2O$ is an impure form of aluminium oxide. In this method the finely powered ore is treated with hot sodium hydroxide solution, the aluminium oxide present in bauxite ore reacts with sodium hydroxide to form water soluble sodium metaaluminate leaving behind the undissolved impurities which is filtered off.

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

sodium
metaaluminate

This filtrate on dilution and stirring gives a precipitate of aluminium hydroxide which is filtered and ignited to get a pure aluminium oxide , which is called Alumina.

$$NaAlO_2 + 2H_2O \longrightarrow Al(OH)_3 + NaOH$$

$$2Al(OH)_3 \quad Al_2O_3 + 3H_2O$$

Alumina

5.Describe Zone- refining method for the purification of metals.

Zone refining: Ultra pure metals and non-metals are obtained by zone refining process. It is also called fractional crystallisation method because this refining is based on the principle that when an impure metal is melted and allowed to solidify, the impurities move away from the solid region and prefer to be distributed in the molten region.

In this method one end of long rod of an element is heated using a small high frequency induction furnace so that a thin cross-section of the metal is melted. When heating unit is moved slowly along the other end of the rod, the molten region solidifies. The impurities are more soluble in the molten liquid than in the solid. Hence, they move towards the molten region.

As the heater unit is moved to the other end of the rod, impurities also move to the same end. This process is repeated, several times until a purity of 99.999% is achieved. The end portion is impure and can be rejected. A noble gas atmosphere is provided during the process in order to prevent the oxidation of the metal.

Fig. Zone Refining

Elements like Germanium, Silicon and Gallium, which are used as semiconductors, are refined by this process.

6.Write a note on physical properties of metals.

Physical properties: Metals have characteristic properties such as high thermal and electrical conductivity, bright lustre, malleability and ductility. The properties of metals can be explained as:

1. **Metallic lustre :** The property of metals due to which they shine is called metallic lustre. The bright lustre of metals is due to their ability to reflect the incident light from their face and the surface acquires a shining appearance, which is known as metallic lustre. Iron, silver, magnesium, aluminium etc. appear as white whereas gold is yellow and copper is reddish brown in appearance.
2. **Metals are solids:** Most of the metals are generally solids at room temperature except mercury, which is liquid at room temperature.
3. **Metals are hard:** Due to strong forces of attraction between the metal atoms, they are generally hard except sodium and potassium which are soft metals. The hardness varies from metal to metal.

   The metals like magnesium (Mg), lead (Pb), aluminium (Al), Iron (Fe), copper (Cu), etc. cannot be cut with a knife, so these are not soft metals.
4. **Metals are malleable and ductile:** Metals can be beaten into sheets (malleability) and drawn into wires (ductility). Metals like Silver, Iron, Copper and Aluminium expand when striked with a hammer. Gold is the most malleable metal. Metals like gold, aluminium, copper silver, *etc.,* can be drawn into thin wires very easily and are, therefore, said to be ductile and this property of metals is called ductility.
5. **Thermal conductivity:** Metals are good conductors of heat. For example iron, silver, gold, aluminium , etc. conduct heat.
6. **Electrical conductivity:** Metals are good conductors of electricity.
7. **Sonorous:** Most of the metals produce ringing sound when they are struck.
8. **Tensile strength:** Metals resist breaking when stretched. This is a measure of their tensile strength. Metals like tungsten have very high tensile strength.
9. **Density:** Metals generally have high densities, *e.g.,* density of gold, mercury and iron is 19.3, 13.6 and 7.6 g/cc. respectively. But some metals, *e.g.,* sodium, potassium, aluminium and magnesium have low densities, lithium has density 0.50 g/cc. It is the lightest metal.

OTHER IMPORTANT QUESTIONS & ANSWERS

I. CHOOSE THE BEST ANSWER:

1. The process of removal of impurities from a crude metal is called
   (a) Concentration (b) Calcination (c) Refining (d) Roasting
2. The impurities present in the ore when mined are called
   (a) flux (b) slag (c) gangue (d) roasting
3. Calcination and roasting are
   (a) different names of the same operation (b) used for the purification of metals
   (c) usually carried in the reverberatory furnance (d) employed for the concentration of the ore
4. Froth floatation process involves the
   (a) treatment of the ore with water and pine oil (b) warming of ore with a stream of water
   (c) pouring of the ore over the belt rotating over magnetic rollers (d) none
5. Which one is true out of the following?
   (a) All ores are minerals but all minerals are not ores
   (b) All minerals are ore but all ores are not minerals (c) Both the above statements are wrong
6. The sulphide ores are generally concentrated by
   (a) gravity separation (b) froth floatation process (c) magnetic separation (d) liquation
7 .For concentration of sulphide ore, following is used
   (a) gravity separation method (b) forth floatation method
   (c) magnetic concentration (d) chemical method
8.The substance mixed in the separation of impurities from ores is
   (a) slag (b) flux (c) catalyst (d) smelter
9 . The process is employed for purification of volatile metals like mercury is ………………
   (a) poling (b) liquation (c) distillation (d) zone refining
10.The process, in which the ore is heated to the extent that it does melt but becomes porous, is called
   (a) roasting (b) smelting (c) calcination (d) decomposition
11.Generally, the extraction of alkali and alkaline earth metals is done by
   (a) electrolytic reduction method (b) reduction with carbon
   (c) alumino thermic process (d) metal displacement method
12.The impurities associated with minerals are called
   (a) Slag (b) Flux (c) Gangue (d) Ore
13.Most of the metals which occur in native state
   (a) Are very reactive (b) have low reactive (c) can form silicates readily (d) are not reactive
14.The process of extracting the metal from its ore is called
   (a) Refining (b) Concentration (c) Leaching (d) Metallurgy
15.The method for the purification of impure metals which is based upon the phenomenon of electrolysis is called
   (a) Electrorefining (b) Hydrometallurgy (c) Poling (d) Liquation
16.Carbon is used as a reducing agent in the extraction of
   (a) Chromium (b) Copper (c) silver (d) Zinc
17. Coke is used in metalurgical process chiefly as:
   (a) flux (b) reducing agent (c) slag (d) oxidizing agent
18. Zone refining has been employed for preparing ultra pure samples of
   (a) Cu (b) Na (c) Ge (d) Zn
19. During roasting of zinc blende, it converts to
   (a) ZnO (b) $ZnSO_4$ (c) $ZnCO_3$ (d) Zn

20. The role of calcination I metallurgical operations is:
   (a) to remove moisture (b) to decompose carbonate
   (c) to drive off organic matter (d) all the above
21. The metal always found in free state is:
   (a) Gold (b) Silver (c) Copper (d) Sodium
22. The process in which lighter earthy particles are freed from the heavier particle by washing with water is called
   (a) Leaching (b) Levigation (c) Liquation (d) Gravity separation
23. Which of the following metals cannot be extracted by carbon reduction process?
   (a) Lead (b) Aluminium (c) Mercury (d) Zinc
24. In electrorefining, the impure metal is made
   (a) cathode (b) anode (c) may be cathode or anode (d) none of these
25. Slag is:
   (a) flux and coke (b) metal and flux (c) coke and metal oxide (d) flux and impurities

26. Roasting is generally done in case of
(a) Oxide ores (b) silicate ores (c) Sulphide ores (d) Carbonate ores
27. In metallurgy, flux is a substance used to convert
(a) insoluble impurities to a fusible mass (b) minerals into silicates
(c) soluble particles into insoluble particles (d) fusible impurities to infusible impurities
28. In aluminothermic process aluminium acts as
(a) oxidizing agent (b) reducing agent (c) flux (d) none of these
29. Which of the following ores cannot be concentrated by froth floatation process?
(a) Bauxite (b) Cinnabar (c) Galena (d) copper pyrite
30. Electrolytic reduction method is used for
(a) highly electropositive metals (b) highly electronegative metals
(c) metalloids (d) lanthanides only
31. Liquation can be used for refining of
(a) Copper (b) Lead (c) Zinc (d) All metals
32. Lighter ore particles are separated from heavier impurities by
(a) Polling (b) gravity separation (c) leaching (d) froth floatation
33. The process of heating the ore in the absence of air is called
(a) Roasting (b) liquation (c) calcination (d) smelting
34. Mond process is used in the refining of
(a) Copper (b) Zirconium (c) Nickel (d) Aluminium
35. In electro refining, impure metal is used as
(a) anode (b) cathode (c) anode or cathode (d) electrolyte
36. Which of the following metals can be refined by liquation?
(a) tin (b) iron (c) magnesium (d) aluminium
37. Which of the following metals is not extracted by reduction with carbon?
(a) Lead (b) Aluminium (c) mercury (d) zinc
38. The metal which is found in liquid state is ………………
(a) potassium (b) copper (c) iron (d) mercury
39. The electropositive element is …………………..
(a) bromine (b) oxygen (c) calcium (d) iodine
40. One of the properties of metals is that …………
(a) they are ductile (b) they do not react with acid
(c) they are poor conductors of heat
(d) they are found at ordinary temperatures in all the three states of matter
41. One of the properties of non –metals is that …………..
(a) they are generally poor conductors of heat (b) they are ductile
(c) they react with acids (d) they are usually solids
42. Aluminium is extracted from the ore ………..
(a) felspar (b) corundum (c) diaspore (d) bauxite
43. A lustrous non –metals is ……….
(a) sodium (b) mercury (c) platinum (d) iodine
44. A metal which is not a good conductor of heat and electricity is …………
(a) copper (b) aluminium (c) iron (d) bismuth
45. A non –metal which is a good conductor of electricity is …………..
(a) oxygen (b) chlorine (c) sulphur (d) carbon in the form of graphite
46. An electropositive element is …………..
(a) sodium (b) calcium (c) potassium (d) All the above
47.The ore which is concentrated by magnetic separation is ……………..
(a) bauxite (b) cinnabar (c) cuprite (d) tin ore
48. Electrolytic refining is hot used to purify …………
(a) copper (b) zinc (c) mercury (d) tin
49. Which of the following reactions represents calcinations?
(a) $CaCO_3 \xrightarrow{heat} CaO + CO_2$ (b) $2PbS + 3O_2 \xrightarrow{heat} 2PbO + 2SO_2$
(c) $ZnO + C \xrightarrow{heat} Zn + CO$ (d) $Cr_2O_3 + 2Al \xrightarrow{heat} 2Cr + Al_2O_3$
50. Which of the following is an example of carbon as a reducing agent
(a) $2CuFeS_2 + 4O_2 \longrightarrow Cu_2S + 2FeO + 3SO_2$ (b) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
(c) $Cu_2S + 2Cu_2O \longrightarrow 6Cu + SO_2$ (d) $3Mn_3O_4 + 8Al \longrightarrow 4M_2O_3 + 9Mn$
51. The changes that occur during roasting could be
(a) Removal of moisture and voltaic matter (b) conversion of sulphide ore into oxide
(c) decomposition of complex ore (d) any one or a combination of 1,2,3

52. Roasting results in the production of metal in the case of
(a) Iron pyrites (b) Galena (c) cinnabar (d) bauxite.
53. The process of heating the ore in the absence of air is called
(a) Roasting (b) liquation (c) calcinations (d) smelting
54. Metals such as copper, silver, etc., are refined by ……………..
(a) zone refining (b) oxidation (c) liquation (d) electrolysis
55. ……………. is extracted by Bessemer process.
(a) Gold (b) Aluminium (c) Iron (d) Silver
56. ………….. is used as semiconductor.
(a) Silicon (b) Lead (c) Graphite (d) Aluminium
57. A ……………… atmosphere is provided in zone refining process to prevent the oxidation of the metal.
(a) charged (b) hot (c) cold (d) inert gas
58. The purity of metal obtained by zone –refining is …………..
(a) 90% (b) 95% (c) 99% (d) 99.999%
59. Ultra pure metals and non-metals are obtained by …………..
(a) poling (b) distillation (c) liquation (d) zone – refining
60. Electro positive metals like sodium and magnesium are extracted from their fused chlorides by ……
(a) reduction (b) electrolysis (c) calcination (d) alumino thermic process
61. ………… is a reducing agent.
(a) carbon (b) carbon monoxide (c) carbon dioxide (d) both (a) and (b)
62. …………… ores are generally calcinated.
(a) sulphide (b) carbonate (c) chloride (d) native
63. The process of heating an ore in absence of excess of air at high temperature is………..
(a) roasting (b) calcination (c) reduction (d) oxidation
64. Zinc sulphide is roasted to get …………..
(a) zinc (b) zinc sulphate (c) zinc oxide (d) zinc carbonate
65. The formula of sodium metaaluminate is …………………
(a) $NaAlO_3$ (b) Na_2AlO_2 (c) $NaAlO_2$ (d) Na_3AlO_6
66. …………. is concentrated by chemical method.
(a) Haematite (b) Magnetite (c) Bauxite (d) Tin stone
67. In froth floatation process, the metallic ore particles get stuck to the froth and rise to the surface as ….
(a) gangue (b) matrix (c) slag (d) scum
68. The formula of calcium hydride is ……………
(a) CaH (b) CaH_2 (c) Ca_2H (d) CaH_4
69. The unwanted impurities such as mud, stones, sand, etc., which are present in the ore are called ….
(a) matrix (b) slag (c) mineral (d) matte
70. ………………. is the gangue in copper pyrite ore.
(a) Calcium silicate (b) Ferrous oxide (c) Ferrous silicate (d) Copper oxide
71. The density of iron is …………… g /cc.
(a) 19.3 (b) 13.6 (c) 7.86 (d) 1
72. The density of mercury is …………….g/cc
(a) 19.3 (b) 13.6 (c) 7.6 (d) 1
73. The density of gold is …………………. g/cc.
(a) 19.3 (b) 13.6 (c) 7.6 (d) 1
74. Non metals are soft except …………..
(a) graphite (b) diamond (c) sulphur (d) iodine
75. …………….. have high tensile strength.
(a) Solids (b) Metals (c) Non-metals (d) Compounds
76. Metals can be drawn into wires. This property is known as …………
(a) malleability (b) ductility (c) conductivity (d) hardness
77. Metals are hard except ……………
(a) sodium (b) potassium (c) aluminium (d) (a) and (b)
78. ………………. are malleable.
(a) Elements (b) Metals (c) Non-metals (d) solids

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

75. (b) 76. (b) 77. (d) 78. (b)

II. Answer the following in one or two sentences:

1.How will you get metals in nature?

Metals like silver, gold, platinum, *etc.,* are less reactive and occur in the native state whereas metals like copper, aluminium, sodium etc., are highly active and occur in the combined state with halogens, sulphur, oxygen, etc. *e.g.,* copper pyrites ($CuFeS_2$), bauxite (Al_2O_3).

2.Name any two soft metals and a hard non-metal.

Soft metals: Sodium & potassium

Hard non-metal: Diamond

3.Why metals are hard?

Due to strong forces of attraction between the metal atoms, they are generally hard except sodium and potassium which are soft metals. The hardness varies from metal to metal.

4.Complete & balance the following:

(a) $2Na + 2H_2O \longrightarrow$

(b) $2K + 2H_2O \longrightarrow$

Sodium and potassium react vigorously with cold water forming their hydroxides and liberating H_2 gas.

(a) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

(b) $2K + 2H_2O \longrightarrow 2KOH + H_2$

5.How will you get metallic hydrides from metals?

Few actives metals like potassium, sodium, calcium reacts with hydrogen to form salts called hydrides.

2Na	+	H_2	$\longrightarrow$	2NaH
Sodium		hydrogen		sodium hydride
Ca	+	H_2	$\longrightarrow$	CaH_2
Calcium		hydrogen		calcium hydride

6. What are minerals?

The various compounds of metals, which occur in nature and are obtained by mining, are called minerals.

7. What are ores?

These are minerals from which metals can be conveniently and economically extracted. *E.g.:* Haematite is the ore of iron because it is used in the extraction of iron.

8. What are the methods adopted for the enrichment of ores?

Methods adopted for the enrichment of ores:

1. Hydraulic washing (or) Gravity separation of method
2. Electro magnetic separation.
3. Froth floatation process.
4. Chemical method.

9. Explain roasting with an example.

It is a process of heating the ore strongly in excess of air. During roasting the voltaile impurities are removed and the ore is changed into oxide. Sulphide ores are generally roasted into oxides. *For example,* zinc sulphide is roasted to get zinc oxide.

$$2ZnS + 3O_2 \longrightarrow 2ZnO + 2SO_2 \uparrow$$

10. Carbon acts as a reducing agent. Explain with example.

Carbon as reducing agent: Moderately reactive metals like iron, zinc, etc., the oxides of these metals are reduced to the metal by heating with carbon in the form of coal, coke, charcoal and carbon monoxide

$$Fe_2O_3 + CO \longrightarrow 2FeO + CO_2\uparrow$$

$$ZnO + CO \longrightarrow Zn + CO_2\uparrow$$

11.What is alumino thermic process?

It is a process of converting a metallic oxide into the corresponding metal by using aluminium.

The oxides of chromium and manganese are reduced by alumino thermic process

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

$$3Mn_3O_4 + 8Al \longrightarrow 4Al_2O_3 + 9Mn$$

12. What is called zone refining? Give its principle.

Zone refining: Ultra pure metals and non-metals are obtained by zone refining process. It is also called fractional crystallisation method because this refining is based on the principle that when an impure metal is melted and allowed to solidify, the impurities move away from the solid region and prefer to be distributed in the molten region.

13. Give the action of dilute acids on metals by giving examples.

Many metals on reaction with dilute acids liberate hydrogen gas.

$$2Na + 2\ HCl \longrightarrow 2NaCl + H_2$$

$$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

Metals like Cu and Ag do not release hydrogen from dilute acids.

14.Define metallurgy.

The process of extraction of pure metals from their ores is called metallurgy.

The method of extraction of metal depends on the nature of the metal and the nature of its ore.

15.Explain the term poling.

The impure metal is melted and the molten metal is stirred with logs of green wood. The impurities are removed either as gases or they get oxidised forming scum over the molten metal.

16.Write a note on distillation.

This process is employed for purification of volatile metals like mercury, zinc and cadmium. The impure metal is heated in a retort and its vapours are separately condensed in a receiver. While the pure metal distills over, the non-volatile impurities are left behind in the retort.

17. Define oxidation.

This method is generally employed in the purification of metals, when the impurities get oxidised more readily than the metal itself. The impure metal is melted and exposed to air in a suitable furnace. The oxides of impurities are formed on the surface. They are removed by skimming.

Eg. Oxygen is blown through molten impure iron in a Bessemer converter.

18. Give examples for active metals.

Some active metals are sodium, potassium, aluminium, etc.

19. Give examples for a few metals which are used in day - today life.

Metals generally used in daily life are iron, copper, tin, lead, silver, nickel and mercury.

20. Give examples for metals which occur in native state.

The metals which occur in native state are silver, gold and platinum.

21. Give example for non –metals.

A few non –metals are hydrogen, carbon, oxygen, Sulphur and phosphorus.

22. Give examples for metals which do not release hydrogen with dilute acids.

Metals like copper and silver do not release hydrogen from dilute acids.

23. Give examples for metals which do not react with water or steam.

Gold, silver and copper do not react with water or steam.

24. Write the equations for the action of sodium and potassium with water.

$$2\ Na + 2H_2O \longrightarrow 2NaOH + H_2$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

25. What happens when PCl_3 and $SiCl_4$ are treated with water? Give equation.

Phosphorus trichloride on hydrolysis gives phosphorus acid. Silicon tetrachloride on hydrolysis gives the corresponding hydroxide.

$$PCl_3 + 3H_2O \longrightarrow 3HCl + H_3PO_3$$

$$SiCl_4 + 4H_2O \longrightarrow Si(OH)_4 + 4HCl$$

26. What happens when aluminium chloride is hydrolysed? Give equation.

Aluminium chloride on hydrolysis gives aluminium hydroxide

$$AlCl_3 + 3H_2O \longrightarrow Al(OH)_3 + 3HCl$$

27. What happens when sodium oxide and sulphur trioxide reacts with water? Give equation.

Sodium oxide on reaction with water gives sodium hydroxide whereas sulphur trioxide gives sulphurous acid.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$SO_3 + H_2O \longrightarrow H_2SO_3$$

28. Give example for a non –metal which is a good conductor of electricity & a liquid metal.

Bromine is a non –metal which is in liquid state. Mercury is a liquid metal.

29. Give examples for metals which are purified by zone –refining.

Germanium, Silicon and Gallium are refined by zone –refining.

30. What is called anode –mud?

The impurities deposited at the bottom of anode during electrolytic refining is known as anode mud.

31. Name the type of ore concentrated by hydraulic washing?

The ores of heavy metallic oxides are concentrated by hydraulic washing.

32. Name the electrode at which pure metal is deposited during electrolysis.

Pure metal is deposited at the cathode.

33. Name the magnetic impurity present in tin stone?

Wolfromite is the magnetic impurity present in tin stone.

34. Name the reducing agent used in Gold Schmidt thermal process?

Aluminium is the reducing agent used in Gold Schmidt thermal process.

35. Name the oil used in froth floatation process?

Pine oil is used in froth floatation process.

36. What type of ores are concentrated by froth floatation process?

Sulphide ores are concentrated by froth floatation process.

37. How do metals react with hydrogen? Give equation.

Active metals like potassium, sodium, calcium react with hydrogen to form salts called hydrides.

$$2Na + H_2 \longrightarrow \underset{\text{Sodium hydride}}{2NaH}$$

$$Ca + H_2 \longrightarrow \underset{\text{Calcium hydride}}{CaH_2}$$

III. Answer in brief:

1.What happens when metals are treated with water? Explain by giving equations.

Metals on reaction with water form metal oxide or metal hydroxide and liberate hydrogen gas.

(i) Sodium and potassium react vigorously with cold water forming their hydroxides and liberating H_2 gas.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

The reactivity of potassium with water is more than that of sodium. This is because, the reaction of potassium with water is so violent that the evolved hydrogen catches fire.

(ii) Magnesium on reaction with water form magnesium oxide and hydrogen.

$$Mg + H_2O \longrightarrow MgO + H_2$$

(iii) Heated iron reacts with steam to give ferroso ferric oxide and hydrogen.

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

(iv) Metals like Au, Ag and Cu do not react even with steam.

2.Explain the term oxidation which respect to metallurgy.

This method is generally employed in the purification of metals, when the impurities get oxidised more readily than the metal itself. The impure metal is melted and exposed to air in a suitable furnace. The oxides of impurities are formed on the surface. They are removed by skimming.

Eg. Oxygen is blown through molten impure iron in a Bessemer converter. The impurities like carbon, sulphur and arsenic are oxidised. The impurities like phosphorus and silicon form oxides which are converted to slag using suitable fluxes. The slag can be removed by skimming.

IV. Answer in detail:

1.How do extracts metals by reduction method? Explain the various steps involved.

The extraction of a metal from one of its ores is essentially a process of reduction (addition of electrons to the metal ion). The various steps involved in reduction process are given below:

1. Roasting is the process of heating the ore strongly in excess of air. During roasting the voltaile impurities are removed and the ore is changed into oxide. Sulphide ores are generally roasted into oxides. *For example,* zinc sulphide is roasted to get zinc oxide.

$$2ZnS + 3O_2 \longrightarrow 2ZnO + 2SO_2 \uparrow$$

2. Calcination is the process of conversion of ore into metal oxide (oxidation) by heating strongly in absence of excess of air at high temperature. During calcination the volatile impurities are removed and the mass become porous. Carbonate and oxide ore are generally calcinated where they lose moisture, carbon dioxide and other volatile impurities.

$$\underset{\text{Malachite}}{CuCO_3 .Cu(OH)_2} \longrightarrow \underset{\text{Copper oxide}}{2CuO} + H_2O + CO_2\uparrow$$

The metal oxides are then reduced to the corresponding metals by using suitable reducing agent. For example

(1) Carbon as reducing agent: Moderately reactive metals like iron, zinc, etc., the oxides of these metals are reduced to the metal by heating with carbon in the form of coal, coke, charcoal and carbon monoxide

$$Fe_2O_3 + CO \longrightarrow 2FeO + CO_2\uparrow$$

$$FeO + CO \longrightarrow Fe + CO_2\uparrow$$

$$ZnO + CO \longrightarrow Zn + CO_2\uparrow$$

2. What is called refining of metals? Discuss any three methods of refining metals.

The process of purification of impure metals by removing metallic and nonmetallic impurities is known as refining of metals.

The impure metals are purified by any one of the following methods.

1. Poling: The impure metal is melted and the molten metal is stirred with logs of green wood. The impurities are removed either as gases or they get oxidised forming scum over the molten metal.

2. Liquation: This method is used for refining those metals, which have low melting point such as tin, lead etc

The impure metal is placed on the sloping hearth of a furnace and gently heated. The metal melts and drain away leaving behind the infusible materials on the hearth.

3. Distillation: This process is employed for purification of volatile metals like mercury, zinc and cadmium. The impure metal is heated in a retort and its vapours are separately condensed in a receiver. While the pure metal distills over, the non-volatile impurities are left behind in the retort.

3. Write short note on Zone Refining.

Ultra pure metals and non-metals are obtained by zone refining process. It is also called fractional crystallisation method because this refining is based on the principle that when an impure metal is melted and allowed to solidify, the impurities move away from the solid region and prefer to be distributed in the molten region.

In this method one end of long rod of an element is heated using a small high frequency induction furnace so that a thin cross-section of the metal is melted. When heating unit is moved slowly along the other end of the rod, the molten region solidifies. The impurities are more soluble in the molten liquid than in the solid. Hence, they move towards the molten region.

As the heater unit is moved to the other end of the rod, impurities also move to the same end. This process is repeated, several times until a purity of 99.999% is achieved. The end portion is impure and can be rejected. A noble gas atmosphere is provided during the process in order to prevent the oxidation of the metal.

Fig. Zone Refining

Elements like Germanium, Silicon and Gallium, which are used as semiconductors, are refined by this process.

4. How do you purify metals by (a) electrolytic refining (b) oxidation methods.

(a) Electrolytic – Refining: This method is most widely used for refining impure metals. Metals such as copper, zinc, tin, nickel etc., are refined electrolytically.

The impure metal to be refined is made, as anode of an electrolytic cell while the cathode is a thin plate of the pure metal. On passing the electric current, the impure metal dissolves, go into solution while insoluble matter settles down at the bottom and is called the anode mud.

(b) Oxidation: This method is generally employed in the purification of metals, when the impurities get oxidised more readily than the metal itself. The impure metal is melted and exposed to air in a suitable furnace. The oxides of impurities are formed on the surface. They are removed by skimming.

Eg. Oxygen is blown through molten impure iron in a Bessemer converter. The impurities like carbon, sulphur and arsenic are oxidised. The impurities like phosphorus and silicon form oxides which are converted to slag using suitable fluxes. The slag can be removed by skimming.

5.What happens when the following are treated with active metals? Explain by giving equations.

(a) O_2; (b) H_2O; (c) HCl;

(a)Metals generally combine with oxygen to from oxides, which are basic in nature.

$$2Zn + O_2 \longrightarrow \underset{\text{Zinc oxide}}{2\,ZnO}$$

$$2\,Mg + O_2 \longrightarrow \underset{\text{Magnesium oxide}}{2\,Mg\,O}$$

(b)Metals on reaction with water form metal oxide or metal hydroxide and liberate hydrogen gas.

(i) Sodium and potassium react vigorously with cold water forming their hydroxides and liberating H_2 gas.

$$2Na + 2\,H_2O \longrightarrow 2NaOH + H_2$$

$$2K + 2\,H_2O \longrightarrow 2\,KOH + H_2$$

The reactivity of potassium with water is more than that of sodium. This is because, the reaction of potassium with water is so violent that the evolved hydrogen catches fire.

(ii) Magnesium on reaction with water form magnesium oxide and hydrogen.

$$Mg + H_2O \longrightarrow MgO + H_2$$

(iii) Heated iron reacts with steam to give ferroso ferric oxide and hydrogen.

$$3\,Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

(c) Many metals on reaction with dilute acids liberate hydrogen gas.

$$2Na + 2\,HCl \longrightarrow 2NaCl + H_2$$

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2$$

Metals like Cu and Ag do not release hydrogen from dilute acids.

6.What happens when the following are treated with active metals? Explain by giving equations.

(i) H_2SO_4; (ii) Cl_2; (iii) H_2?

(i) Metals like magnesium reacts with dilute sulphuric acid forming magnesium sulphate by the liberation of hydrogen gas.

$$Mg + H_2SO_4 \longrightarrow MgSO_4 + H_2$$

(ii) Metals reacts directly with halogens like chlorine to form electrovalent compounds which have the properties of salts.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

$$Ca + Cl_2 \longrightarrow CaCl_2$$

(iii) Few actives metals like potassium, sodium, calcium reacts with hydrogen to form salts called hydrides.

$$\underset{\text{Sodium}}{2Na} + \underset{\text{hydrogen}}{H_2} \longrightarrow \underset{\text{sodium hydride}}{2NaH}$$

$$\underset{\text{Calcium}}{Ca} + \underset{\text{hydrogen}}{H_2} \longrightarrow \underset{\text{calcium hydride}}{CaH_2}$$

6 - METALS

INTRODUCTION:

A careful observation of the periodic table reveals that, it is dominated by metallic elements. In fact, there are 80 metallic elements out of 115 elements known so far.

Metals can be divided into two categories, namely lighter metals and heavier metals. Metals, which possess the density below 4 are treated as lighter metals. The rest are heavier metals. Thus the lighter metals include alkali metals (Li, Na, K, Rb, Cs), alkaline earth metals (Ca, Sr, Ba, Mg) and aluminium, while the heavier metals include base metals (Pb, Su, Cd, Fe, Co, Ni, Cr, Mn etc.) and noble metals (Cu, Ag, Au, Hg, Pt etc.).

Lighter metals are mostly prepared by the electrolytic reduction on their compounds. Heavier metals are generally prepared by chemical reduction of their oxides and sulphides. Lighter metals are chemically reactive. Heavier metals are relatively much less reactive.

Metals and their alloys are the backbone of all engineering projects and products. Compounds of metals also find many applications in our daily life.

IMPORTANT TERM & DEFINTIONS:

The important minerals of aluminium are: 1.Bauxite ($Al_2O_3 . 2H_2O$) 2.Cryolite (Na_3AlF_6)
3. Corundum (Al_2O_3)

The extraction of aluminium requires three stages:
(a) Purification of bauxite (b) Electrolytic reduction (c) Refining of aluminium
Aluminium is a self-protecting metal because the oxide film on aluminium protects it from further attack of air.

The minerals of copper are:
1. Copper pyrites ($CuFeS_2$) 2. Copper glance (Cu_2S.) 3. Cuprite (Cu_2O.)
4. Malachite ($Cu(OH)_2 . CuCO_3$.) 5. Azurite ($2CuCO_3 . Cu(OH)_2$.)

Concentration of Ore: Copper pyrites being a sulphide ore, is concentrated (dressed or enriched) by the **froth- floatation** process.

Roasting: The concentrated copper pyrites ore is roasted in air in a blast furnace.

Conversion to Metal: When a good amount of copper sulphide has been converted into copper oxide, then after some time, the supply of air for roasting is stopped. In the absence of air, Copper oxide formed above reacts with the remaining copper sulphide to form copper metal:

Bessemerization: The Process in which copper oxide reacts with copper sulphide to form copper metal is called " bessemerization".

Electrolytic refining: Impure copper metal is refined by electrolysis method called electrolytic refining.

The minerals of iron are: 1.Haematite(Fe_2O_3) 2.Magnetite (Fe_3O_4) 3.Siderite($FeCO_3$)
4.Limonite($2Fe_2O_3 . 3H_2O$)

Concentration (levigation): The powdered ore is washed with a stream of water whereby the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

Calcination: The Concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. During roasting (a) moisture is driven out and (b) impurities like sulphur, arsenic, phosphorus etc are oxidised off.

Smelting: The roasted ore is mixed with limestone and coke and heated in a blast furnace in order to reduce the iron oxide to the metal.

Tuyers: These are *small pipes* (tuyers) through which a blast of hot air is admitted and a *slaghole* through which slag can be withdrawn.

Types of zones: The various types of temperature zones are (a) combustion zone (b) fusion zone
(c) slag formation zone (d) reduction zone

Three commercial forms of iron: (a) Cast iron (b) Wrought iron and (c) steel.

Cast iron: It is the most impure form of iron containing 2-4.5% of carbon. It is very hard and brittle.

Wrought iron: It is the purest form of iron obtained by the removal of carbon almost completely. It contains less than 0.25% carbon.

Steel: Steel is an alloy of iron with 0.25% to 2% carbon. The percentage of carbon in steel is intermediate between that in wrought iron and in cast iron.

Different types of steel: Based on carbon content, there are three types of steel.

Mild steel: It has the least carbon content; 0.1 to 0.15%. It is used for making wires and sheets.

Medium steel: It contains 0.2 to 0.5% carbon. It is harder than mild steel and is used for constructing rails, wheels etc.

Hard steel: This type of steel contains 0.5 to 1.5% carbon. It is very hard and used for making machine parts.

Special steels or alloy steels: Steel mixed with small amount of nickel, cobalt, chromium, tungsten, molybdenum, manganese, etc., acquires special properties. Such products are called special steels or alloy steels. Some important alloy steels are described.

Corrosion: It may be defined as the slow and steady destruction of a metal or alloy by the environment.

In case of iron, corrosion is called rusting. Rust is a hydrated ferric oxide represented as $Fe_2O_3.H_2O$.

Methods Of Preventing Corrosion

Coating with paints: Metal surfaces coated with paint which keep it out of contact with air, moisture, etc., till the paint layer develops cracks.

Coating with oils and greases: By applying film of oil and grease on the surface of the iron tools and machinery, the rusting of iron can be prevented since it keeps the metal surface away from moisture, oxygen and carbon dioxide.

Alloying: Some metals, when mixed with other metals, become resistant to corrosion. Stainless steel is an alloy of iron, which does not undergo corrosion easily.

Galvanisation: This process involves the coating of zinc on iron sheets to prevent rusting. Galvanised iron is used to make buckets, boxes, utensiles etc., and other commonly used articles.

Tinning: This process involves the coating of tin (with molten tin) on cooking vessels made of copper and brass.

Anodizing: In this process metals like aluminium, copper, etc. are coated electrically with a thin and strong film of their oxides which protects them from rusting. Articles such as soap cases, handles, doorknobs, etc., are commonly anodized aluminium articles.

Electroplating: It is a process of depositing stable metal (gold, silver) over a base metal (copper, iron) Iron can be coated with copper by electro deposition from a solution of copper sulphate.

SELF EVALUATION (T.B. PAGE 108)

I. Choose the correct answer.

1. The principal ore of aluminium is ....
(a) Bauxite (b) Siderite (c) Azurite (d) Malachite
2. Bauxite ore is purified by ....
(a) Alumino thermic process (b) Baeyer's process
(c) Hall's process (d) Froth flotation process
3. Alnico is an alloy of ....
(a) Al + Cu + Mn + Mg (b) Al + Fe + Ni + Co
(c) Al + Mg (d) Al + Cu
4. Bronze is an alloy of ....
(a) Zinc (b) Copper (c) Nickel (d) Iron
5. The possible valancies of copper are ....
(a) 1, 2 (b) 2, 3 (c) 3 (d) 3, 4
6. Copper reacts with dil. HNO_3, it gives mainly ....
(a) Nitric oxide (b) Nitrogen dioxide
(c) Nitrous oxide (d) Nitrous acid
7. During extraction of iron the flux used is ....
(a) Silica (b) Calcium silicate
(c) Lime stone (d) Coke
8. Which one of the following statement is not correct?
(a) Wrought iron is the purest form of steel.
(b) Pig iron is the impure form of steel.
(c) Hardness of steel decrease with increase in carbon content.
(d) Wrought iron is the impure form of steel
9. The objects like pipes, stoves and hot water radiators are prepared by using ....
(a) Wrought iron (b) cast iron
(c) pig iron (d) steel
10. The process of coating zinc over iron sheets is known as ....
(a) Galvanisation (b) Tinning
(c) Anodizing (d) Alloying

Answers:

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

II. Answer the following in One or Two sentences (T.B Page 102)

1.What is the action of water on aluminium?

Pure water has almost no action on aluminium in cold. Salt water (*e.g.,* sea water) corrodes it rapidly especially when it is hot. It decomposes boiling water liberating hydrogen.

$$2Al + 6H_2O \longrightarrow 2Al(OH)_3 + 3H_2\uparrow$$

2.Name the constituents of Y-alloy.

Aluminium, copper, nickel magnesium, silicon, iron.

3.What are the important minerals of copper?

The minerals of copper are:

1. Copper pyrites ($CuFeS_2$)
2. Copper glance (Cu_2S)
3. Cuprite (Cu_2O)
4. Malachite ($Cu(OH)_2.CuCO_3$)
5. Azurite ($2CuCO_3.Cu(OH)_2$)

4.State four uses of copper.

1. Due to the electrical conductivity, it is used for making electric cables and other electrical goods.
2. Due to its thermal conductivity, it is used in making utensils, boilers and calorimeters.
3. It is used in electroplating, electrotyping and making coins.
4. It is used in the manufacture of insecticides, pesticides pigments etc.

5.Define the term galvanization.

This process involves the coating of zinc on iron sheets to prevent rusting. The layer on the surface of iron, when comes in contact with moisture, oxygen and carbon dioxide in air, a protective invisible thin layer of basic zinc carbonate $ZnCO_3$, $Zn(OH)_2$ is formed. Hence, the galvanised iron sheets do not lose their lustre and also tends to protect it from further corrosion. Galvanised iron is used to make buckets, boxes, utensiles *etc.,* and other commonly used articles.

III. Answer in brief (T.B. Page 109)

1.Write the uses of aluminium.

1. Aluminium is used in making electric cables and transmission wires because of its high electrical conductivity.
2. It is used in utensils because of its high thermal conductivity
3. Aluminium alloys are used in aeroplane parts, surgical instruments etc. The alloys of aluminium are known for their lightness and lustrous properties.
4. Used as a reducing agent in the extraction of chromium and manganese by Gold Schmidt aluminothermic process.
5. Since it is unattacked by concentrated nitric acid, it is used in chemical plants and also for transporting nitric acid.
6. Aluminium powder is used in making silver paints, fire works, flash light powders and thermic welding.
7. A mixture of aluminium powder and ammonium nitrate called ammonal is used in explosives.

2.Explain the process “Bessemerization’.

The Process in which copper oxide reacts with copper sulphide to form copper metal is called “bessemerization”.

In the absence of air, Copper oxide formed roasting reacts with the remaining copper sulphide to form copper metal.

$$CuS + 2CuO \longrightarrow \underset{\text{Copper metal}}{3Cu} + \underset{\text{Sulphur dioxide}}{SO_2}$$

The copper metal formed here is in the molten state. We have just seen that when copper oxide reacts with copper sulphide to form copper metal, then sulphur dioxide gas is also formed.

When this sulphur dioxide gas comes out through the molten copper, then a kind of blisters develops on the surface of copper metal. And because of this, it is called blister copper. From this impure blister copper, pure copper metal is obtained by the process of electrorefining (electrolytic refining).

3.What is the action of acids on iron?

(1) Reaction with hydrochloric acid: Iron liberates hydrogen on reaction with dilute or con. Hydrochloric acid.

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$$

(2) Reaction with sulphuric acid: Hydrogen is liberated on reaction with dilute sulphuric acid and sulphur dioxide is formed with hot concentrated sulphuric acid.

$$\underset{(dil)}{Fe + H_2SO_4} \longrightarrow FeSO_4 + H_2\uparrow$$

$$\underset{(conc.)}{Fe + H_2SO_4} \quad FeSO_4 + 2H_2O + SO_2\uparrow$$

(3) Reaction with nitric acid: Iron reacts with dilute nitric acid forming iron (II) nitrate or ferrous nitrate and ammonium nitrate

$$4Fe + 10HNO_3 \longrightarrow \underset{\text{Ferrous nitrate}}{4Fe(NO_3)_2} + \underset{\text{Ammonium nitrate}}{NH_4NO_3} + 3H_2O$$

4.Explain the three commercial form of iron.

There are three important forms of iron, which are classified mainly on the basis of their carbon content.

(i) Cast iron (ii) Wrought iron and (iii) Steel

(i) Cast iron: It is the most impure form of iron containing 2 - 4.5% of carbon. It is very hard and brittle.

(ii) Wrought iron: It is the purest form of iron obtained by the removal of carbon almost completely. It contains less than 0.25% carbon.

(iii) Steel: Steel is an alloy of iron with 0.25% to 2% carbon. The percentage of carbon in steel is intermediate between that in wrought iron and in cast iron. Hence the properties of steel are also intermediate between those of cast iron and wrought iron. Steel is malleable and ductile. Steel containing other elements such as chromium, nickel, tungsten, vanadium, silicon, manganese are called alloy steels.

5.Give the action of aluminium with dilute and concentrated sulphuric acid.

1. **Reaction with dilute sulphuric acid:** Aluminium reacts with dil. H_2SO_4 to form aluminium sulphate and hydrogen.

$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2\uparrow$$

2. **Reaction with concentrated sulphuric acid:** Aluminium reacts with hot concentrated sulphuric acid to form aluminium sulphate and sulphur dioxide.

$$2Al + 6H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3SO_2 + 6H_2O$$

IV. Answer in detail (T.B. Page 109)

1.Explain the extraction of aluminium from its ore?

The important minerals of aluminium are:

1. Bauxite ($Al_2O_3.2H_2O$) 2. Cryolite (Na_3AlF_6) 3. Corundum (Al_2O_3)

Extraction: The extraction of aluminium requires two stages:

(a) Purification of bauxite (b) Electrolytic reduction

(a). Purification of bauxite by Baeyer's process: The bauxite ore is powdered and roasted to convert any ferrous oxide impurity into ferric oxide. **It is then digested with a concentrated solution of caustic soda (NaOH) at 150°C in an autoclave under pressure**. Aluminium oxide present in the ore dissolves to give sodium meta aluminate while the impurities are left behind.

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

Ferric oxide and silica impurities are removed by filtration. Sodium metaaluminate solution is diluted with water and then agitated for a long time. Freshly prepared aluminium hydroxide is added to act as a seeding agent. The whole of sodium metaaluminate precipitates as aluminium hydroxide due to hydrolysis.

$$NaAlO_2 + 2H_2O \longrightarrow Al(OH)_3\downarrow + NaOH$$

The precipitate of aluminium hydroxide is filtered, washed and ignited (*i.e.* strongly heated) to form pure alumina.

$$2Al(OH)_3 \longrightarrow Al_2O_3 + 3H_2O$$

(b). Electrolytic reduction of alumina by Hall's electrolytic process:

1. Pure alumina is dissolved in molten cryolite and mixed with a little amount of fluorspar, which lowers the melting point of the electrolyte.
2. The molten electrolyte is taken in a rectangular cast iron tank lined with a layer of graphite.
3. The iron tank lined with graphite acts as the cathode.
4. A bunch of graphite rods are suspended in the molten electrolyte to act as the anode.
5. The electrolysis is carried out at 900 – 950°C.During electrolysis, molten aluminium is produced at the cathode and oxygen gas is evolved at the anode.

$$\underset{\text{Aluminium oxide}}{Al_2O_3} \longrightarrow \underset{\text{Aluminium ions}}{2Al^{3+}} + \underset{\text{Oxide ions}}{3O^{2-}}$$

At cathode $2Al^{3+} + 6e^- \longrightarrow 2Al\ (l)$

At anode $6O^{2-} \longrightarrow 3O_2(g) + 12e^-$

The oxygen gas evolved at the anode reacts with graphite anodes and forms carbon dioxide gas.

$$C\ (s) + O_2(g) \longrightarrow CO_2(g)$$

The anodes, thus burn away. Therefore, they must be replaced from time to time.

Fig. Electrolitic cell for extraction of Al

2.How does copper obtained from copper pyrites?

The various steps involved in the extraction of copper metal from copper pyrites ore or sulphide ore are given below:

1. Concentration of Ore:

1. Copper pyrites being a sulphide ore, is concentrated (dressed or enriched) by the **froth- floatation** process.
2. The sulphide ore is powdered and put in a tank full of water containing some pine oil. The mixture in the tank is agitated by blowing in air due to which froth is formed.
3. The sulphide ore particles stick to the froth bubbles and rise to the surface with froth. Gangue particle (impurities) do not stick to the froth and remain at the bottom of the tank.
4. The froth is separated and kept aside for some time. On keeping, foam settles down and concentrated sulphide ore is obtained.

2. Roasting: The concentrated copper pyrites ore is roasted in air in a blast furnace. On roasting , a part of copper sulphide (CuS) present in copper pyrites is oxidised to copper oxide (CuO), and sulphur dioxide gas is evolved:

$$2\,CuS + 3O_2 \longrightarrow \underset{\text{Copper oxide}}{2CuO} + \underset{\text{Sulphur dioxide}}{2SO_2}$$

3. Conversion to Metal: When a good amount of copper sulphide has been converted into copper oxide, then after some time, the supply of air for roasting is stopped. In the absence of air, Copper oxide formed above reacts with the remaining copper sulphide to form copper metal.

$$CuS + 2CuO \longrightarrow \underset{\text{Copper metal}}{3Cu} + \underset{\text{Sulphur dioxide}}{SO_2}$$

The Process in which copper oxide reacts with copper sulphide to form copper metal is called **"bessemerization"**. The copper metal formed here is in the molten state. We have just seen that when copper oxide reacts with copper sulphide to form copper metal, then sulphur dioxide gas is also formed.

When this sulphur dioxide gas comes out through the molten copper, then a kind of blisters develops on the surface of copper metal. And because of this, it is called blister copper. From this impure blister copper, pure copper metal is obtained by the process of electrorefining (electrolytic refining).

4. Electrolytic refining of copper: Impure copper metal is refined by electrolysis method called electrolytic refining. For the electrolytic refining of copper.

Fig. Electro refining

(i) A thick plate of impure copper metal is made anode (+ve electrode)

(ii) A thin plate of pure copper metal is made cathode (–ve electrode)

(iii) Copper sulphate solution (acidified with sulphuric acid) is taken as electrolyte

When a current of electricity is passed between the electrodes, copper is transferred from the anode to the cathode. The impurities in the crude copper collects below the anode as anode mud.

At Cathode: $\underset{\text{(from electrolyte)}}{Cu^{2+}} \xrightarrow[\text{(Reduction)}]{} \underset{\text{Copper atom (Deposits on cathode)}}{Cu}$

At Anode: $\underset{\text{Copper (from impure anode)}}{Cu} \xrightarrow[\text{(Oxidation)}]{} \underset{\text{Copper ion}}{Cu^{2+}}$

3.Write a note on chemical properties of iron.

1. Action of damp air: In moist air, iron is oxidized to form rust

$$4Fe + 3O_2 + 3H_2O \longrightarrow \underbrace{Fe_2O_3 + 2Fe(OH)_3}_{\text{Rust}}$$

2. Action with water: Iron is unaffected by pure cold water. Hydrogen is liberated when steam is passed over red-hot iron.

$$3Fe + 4H_2O \longrightarrow \underset{\text{Ferroso ferric oxide}}{Fe_3O_4} + 4H_2\uparrow$$

3. Action with chlorine: It forms ferric chloride when heated with chlorine

$$2Fe + 3Cl_2 \longrightarrow \underset{\text{Ferric chloride}}{2\,FeCl_3}$$

4. Action of acids:

(i) Reaction with hydrochloric acid: Iron liberates hydrogen on reaction with dilute or con. Hydrochloric acid.

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2\uparrow$$

(ii) Reaction with sulphuric acid: Hydrogen is liberated on reaction with dilute sulphuric acid and sulphur dioxide is formed with hot concentrated sulphuric acid.

$$Fe + \underset{\text{(dil)}}{H_2SO_4} \longrightarrow FeSO_4 + H_2\uparrow$$

$$Fe + \underset{\text{(conc.)}}{H_2SO_4} \qquad FeSO_4 + 2H_2O + SO_2\uparrow$$

(iii) Reaction with nitric acid: Iron reacts with dilute nitric acid forming iron (II) nitrate or ferrous nitrate and ammonium nitrate

$$4Fe + 10HNO_3 \longrightarrow \underset{\text{Ferrous nitrate}}{4Fe(NO_3)_2} + \underset{\text{Ammonium nitrate}}{NH_4NO_3} + 3H_2O$$

4.Describe the extraction of iron from its ore.

The minerals of iron are:

1. Haematite (Fe_2O_3) 2. Magnetite (Fe_3O_4) 3. Siderite ($FeCO_3$) 4. Limonite ($2Fe_2O_3.3H_2O$)

Extraction: The various steps involved during extraction is given below:

1. Concentration: The powdered ore is washed with a stream of water whereby the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

2. Calcination: The Concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. During roasting *(a)* moisture is driven out and *(b)* impurities like sulphur, arsenic, phosphorus etc are oxidised off.

$$S + O_2 \longrightarrow SO_2$$

$$4P + 5O_2 \longrightarrow P_4O_{10}$$

$$4As + 5O_2 \longrightarrow 2\ As_2O_5$$

3. Smelting: The roasted ore is mixed with limestone and coke and heated in a **blast furnace** in order to reduce the iron oxide to the metal.

Blast furnace is made of steel, lined inside with fire resistant bricks. It has a cup and cone arrangement for the introduction of charge at the top, and at the base it has a *tapping hole* through which molten iron can be withdrawn. These are *small pipes* (tuyers) through which a blast of hot air is admitted and a *slaghole* through which slag can be withdrawn.

A mixture of roasted ore, limestone (flux) and coke is fed into the furnace from the top by means of the cup and cone arrangement and a hot blast of air is sent through the tuyers near the base. As a result of the different temperatures attained at different levels in the furnace, the iron oxide gets reduced to iron.

The various reactions at the temperature zones are:

(i) Coke burns at the base to produce CO_2 which rises up:

$$C + O_2 \longrightarrow CO_2 + Heat$$

This reaction is exothermic and hence the temperature here is about 1775K. This region is called ***combustion zone.***

(ii) As CO_2 rises up it is reduced to CO with the coke:

$$CO_2 + C \longrightarrow 2CO - Heat$$

This reaction is endothermic and in this region the temperature falls to 1475 –1575K. Fe_2O_3, if present, gets reduced to iron by hot coke and the spongy iron produced in the upper region gets melted.

$\therefore$ This region is called ***fusion zone.***

Fig. Blast Furnance

(iii) In the middle portion of the furnace where the temperature is about 1075 to 1275 K. Limestone decomposes to produce CaO which combines with silica (impurity) to **form slag**.

$$CaCO_3 \longrightarrow CaO + CO_2$$
$$CaO + SiO_2 \longrightarrow \underset{\text{Slag}}{CaSiO_3}$$

This region is called ***slag formation*** zone.

(iv) Near the top of the furnace where the temperature is about 875K, the oxides of iron are reduced to iron by CO.

$$Fe_2O_3 + CO \longrightarrow 2FeO + CO_2$$
$$FeO + CO \longrightarrow Fe + CO_2$$

This region is called ***reduction zone*** and iron formed moves down and melts in the fusion zone.

The molten slag is removed first and then the molten metal. The iron thus formed is called ***pig iron***. The pig iron after remelting in a vertical furnace is known as cast iron because it can be cast into moulds.

5.Discuss the various methods followed for preventing corrosion.

Corrosion of metals can be prevented if metal surface is not allowed to come in contact with moisture, oxygen and carbon dioxide. **This can be achieved by the following methods:**

1. **Coating with paints:** Metal surfaces coated with paint which keep it out of contact with air, moisture, etc., till the paint layer develops cracks.
2. **Coating with oils and greases:** By applying film of oil and grease on the surface of the iron tools and machinery, the rusting of iron can be prevented since it keeps the metal surface away from moisture, oxygen and carbon dioxide.
3. **Alloying:** Some metals, when mixed with other metals, become resistant to corrosion. Stainless steel is an alloy of iron, which does not undergo corrosion easily.
4. **Galvanisation:** This process involves the coating of zinc on iron sheets to prevent rusting. The layer on the surface of iron, when comes in contact with moisture, oxygen and carbon dioxide in air, a protective invisible thin layer of basic zinc carbonate $ZnCO_3$, $Zn(OH)_2$ is formed. Hence, the galvanised iron sheets do not lose their lustre and also tends to protect it from further corrosion. Galvanised iron is used to make buckets, boxes, utensiles *etc.,* and other commonly used articles.
5. **Tinning:** This process involves the coating of tin (with molten tin) on cooking vessels made of copper and brass. These vessels get a greenish coating due to moist air or due to foodstuffs. The greenish coating (due to corrosion) is poisonous.

   Therefore, vessels are coated with tin. In case of tin coating on iron, the film will be effective as long as it is unscratched. When scratches occur at the coating surface both the metals are exposed to oxygen and rusting of iron occurs.
6. **Anodizing:** In this process metals like aluminium, copper, *etc.,* are coated

   electrically with a thin and strong film of their oxides which protects them from rusting. Articles such as soap cases, handles, door knobs, *etc.,* are commonly anodized aluminium articles.
7. **Electroplating:** It is a process of depositing stable metal (gold, silver) over a base metal (copper, iron). Iron can be coated with copper by electro deposition from a solution of copper sulphate. Spoons, handles of bicycles, taps and many other articles are electroplated. Many gift articles, medals, etc., are gold plated or silver plated electrolytically.

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1. Electrolytic reduction of alumina to aluminium by Hall – Heroult's process is carried out in the presence of
   (a) Na*Cl* (b) fluorite
   (c) cryolite which forms a melt with a lower melting point.
   (d) cryolite which forms a melt with a high melting point.
2. In the Hall's process aluminium is obtained by
   (a) electrolysis of aluminium oxide in molten cryolite
   (b) heating aluminium oxide in an atmosphere of CO_2
   (c) heating aluminium oxide with carbon (d) heating aluminium oxide in an atmosphere of N_2
3. The electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ represets
   (a) Na (b) Al (c) Fe (d) Cu

4. Element X is a lustrous solid that conducts electricity. It reacts with hydrochloric acid also with NaOH solution. Identify X.
(a) Na (b) Cu (c) Fe (d) Al
5. The common oxidation state of aluminium in its compound is
(a) +2 (b) + 1 (c) +3 (d) –4
6. In the Bayer's process of concentration of bauxite which of the following remains insoluble?
(a) alumina (b) silica (c) alumina and silica (d) ferric oxide
7. In which of the following reactions aluminium acts as a reducing agent?
(a) $2Al + 2NaOH + 2H_2O \longrightarrow 2NaAlO_2 + 3H_2$ (b) $4Al + 3H_2SO_4 \longrightarrow 2Al_2(SO_4)_3 + 8H_2$
(c) $Cr_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Cr$ (d) all of 1,2,3
8. Aluminothermic process is suitable for the reduction of oxide of
(a) aluminium (b) chromium (c) copper (d) none of above
9. One of the following alloys is used in making electromagnets. Which is it?
(a) duralumin (b) magnalium (c) y – alloy (d) alnico
10.Which is most suitable for storing con.nitric acid?
(a) copper container (b) aluminium container (c) tin container (d) lead container
11. Which of the following metals is rendered possive by conc HNO_3?
(a) Fe (b) Cr (c) Al (d) all of 1,2,3
12. In the extraction of iron, the charge is a mixture of heamatite, coke and limestone. The effective reducing agent for the oxide
(a) carbon (b) limestone (c) carbon monoxide (d) carbon dioxide
13. The chemical process involved in the production of steel from haematite involves
(a) reduction (b) oxidation
(c) reduction followed by oxidation (d) oxidation followed by reduction
14. Which of the following is a wrong statement about cast iron?
(a) it contains about 4.5% C (b) it is also called pig iron
(c) it contracts on cooling (d) it is corrosion resistant
15. The oxides of iron are reduced to iron in .... Zone.
(a) fusion (b) slag formation (c) reduction (d) oxidation
16. Which of the following reducing agent is used in the extraction of copper
(a) carbon (b) carbon monoxide (c) aluminium (d) none of the above
17. Which of the following is the carbonate ore of copper?
(a) malachite (b) azurite (c) cuprite (d) malachite and azurite
18. Which of the following is not a copper ore?
(a) cuprite (b) cinnabar (c) malachite (d) azurite
19. The oxidation state (s) shown by copper is
(a) +1 (b) +2 (c) +1 and +2 (d) +1 and –1
20. The most abundant element in earth's crust is
(a) Silicon (b) Aluminium (c) nitrogen (d) oxygen
21. Which of the following is formed as Thomas slag during steel manufactured?
(a) $Mn SiO_3$ (b) $CaSiaO_3$ (c) $Ca_3 (PO_4)_2$ (d) $Na_3 ASO_3$
22. Besides iron and C, stainless steel contains
(a) Al and Ni (b) Cr and Ni (c) Cr and Co (d) Co and Ni
23. Purest form of iron is
(a) Cast iron (b) Wrought iron (c) Steel (d) Stainless steel
24. The variety of steel containing least amount of carbon is known as
(a) Mild steel (b) hard steel (c) high carbon steel (d) Stainless steel.
25. Which of the following process would give soft variety of steel?
(a) Quenching (b) tempering (c) dodging (d) anealing
26. The least pure commercial form of iron is
(a) Cast iron (b) Wrought iron (c) Steel (d) haematite.
27. The temperature in the slag formation zone in the blast furnace is
(a) 1775-2000 k (b) 875-1075 K (c) 1075-1475 K (d) 11075-975 K.
28. The steel used for making rock drills and safes is
(a) Stainless steel (b) nickel steel (c) invar (d) manganese steel.
29. When copper is treated with conc. HNO_3, the gas evolved is
(a) H_2 (b) NH_3 (c) NO (d) NO_2
30. The chemical Composition of rust is
(a) $Fe_2O_3.xH_2O$ (b) FeO_4. XH_2O (c) $FeO.xH_2O$ (d) Fe_3O_4.
31. The most abundant transition element is
(a) Iron (b) Copper (c) Aluminium (d) Zinc.

32. The most pure form of iron is
(a) White cast-iron (b) Grey cast iron (c) Wrought iron (d) Steel.
33. Brass is an alloy of
(a) Silver (b) Copper (c) tin (d) zinc
34. The metal used for galvanization of iron sheets is
(a) Nickel (b) Zinc (c) Chromium (d) Aluminium.
35. The slag obtained during the extraction of copper pyrites is composed mainly of
(a) Cu_2S (b) $FeSiO_3$ (c) $CuSiO_3$ (d) SiO_2.
36. Galvanization of iron denotes coating with
(a) Al (b) Sn (c) Cd (d) Zn
37. Brass is an alloy of
(a) Al (b) Cu (c) Ni (d) Zn.
38. Th iron obtained from blast furnace is
(a) Pig iron (b) Wrought iron (c) Soft iron (d) Steel.
39. In the extraction of iron, slag is produced which is
(a) CO (b) $FeSiO_3$ (c) $MgSiO_3$ (d) $CaSiO_3$
40. Which following elements does not show variable valancy?
(a) Copper (b) Iron (c) Zinc (d) Aluminium
41. Invar, an alloy of Fe and Ni is used in watches and meter scales. Its characteristic property is
(a) Small coefficient of expansion (b) Resistance to expansion
(c) Hardness and elasticity (d) Resistance to water.
42. Iron can be coated with copper by electro deposition from a solution of ………………
(a) $CuSO_4$ (b) CuS (c) $CuCO_3$ (d) H_2SO_4
43. The formula of rust is ……………..
(a) Fe_2O_3 (b) $Fe_2O_3H_2O$ (c) $FeCO_3$ (d) FeS
44. The slow and steady destruction of a metal is known as …………..
(a) decomposition (b) splitting (c) corrosion (d) galvanisation
45. An alloy of iron, chromium and nickel is ………….
(a) invar (b) stainless steel (c) german silver (d) chrome steel
46. Helmets are made of ……………
(a) manganese steel (b) chrome vanadium steel (c) hard steel (d) mild steel
47. Pendulums are made of ……………
(a) stainless steel (b) invar (c) tungsten steel (d) brass
48. Invar contains ……………. percentage of nickel.
(a) 10 (b) 25 (c) 36 (d) 64
49. Machines parts are made of …………..
(a) mild steel (b) pig iron (c) medium steel (d) hard steel
50. Wheels are made of ……………
(a) mild steel (b) medium steel (c) hard steel (d) stainless steel
51. The valency of iron 26 is ………………
(a) 1 (b) 2 (c) 3 (d) (b) and (c)
52. …………… is the second most abundant metal in the earth's crust.
(a) Aluminium (b) copper (c) iron (d) silicon
53. ……………… have specialized in the metallurgy of iron and manufactured steel.
(a) Indians (b) Americans (c) south Africans (d) British Scientists
54. ……………… near Kutabminar of Delhi speaks of the glory of our skill in this art.
(a) Taj Mahal (b) Red Fort (c) Iron gate (d) Ashoka's pillar
55. …………… is the principal ore if iron.
(a) Haematite (b) Siderite (c) Limonite (d) Iron pyrites
56. Haematite is concentrated by ………………. Process.
(a) gravity separation (b) froth floatation (c) electromagnetic (d) chemical
57. The concentrated iron ore is strongly heated in a limited supply of air in a …………..
(a) blast furnace (b) reverberatory furnace (c) Bessemer converter (d) open hearth furnace
58. Heating of concentrated ore in a limited supply of air is known as ………..
(a) roasting (b) bessemerisation (c) smelting (d) calcination

59. ……………… is used in the manufacture of insecticides, pesticides, pigments, etc.
(a) Cu (b) Ag (c) Au (d) Al

60. Copper reacts with Con. HNO_3 to produce ………………..
(a) H_2 (b) NO (c) NO_2 (d) O_2

61. Copper reacts with Con. H_2SO_4 to produce ……………..
(a) H_2 (b) SO_2 (c) O_2 (d) H_2S

62. ……………. does not react with water under any condition.
(a) iron (b) aluminium (c) copper (d) sodium

63. The density of copper is …………….. g/cc
(a) 8.94 (b) 9.84 (c) 4.89 (d) 7.6

64. The m.pt of copper is ………………..
(a) 1083 K (b) 1083°C (c) 923°C (d) 923 K

65. …………….. is used in chemical plants and also for transporting nitric acid.
(a) Iron (b) copper (c) Aluminium (d) Lead

66. Aluminium is unattacked by ………………
(a) HCl (b) dil. H_2SO_4 (c) Con. H_2SO_4 (d) HNO_3

67. ……………… is used in making silver paints.
(a) Alumina (b) Aluminium powder (c) Silver (d) Silver salts

68. A mixture of aluminium powder and ammonium nitrate is called ……….
(a) amatol (b) ammonal (c) dynamite (d) TNT

69. The alloy used in air craft parts is ………………
(a) duralumin (b) magnalium (c) Y – alloy (d) alnico

70. An alloy of aluminium and magnesium is called …………….
(a) duralumin (b) magnalium (c) Y – alloy (d) alnico

71. The temperature of the electrolytic bath in Hall's process is …………..
(a) 900 – 950°C (b) 200 – 250°C (c) 900 – 950K (d) above 1000°C

72. The density of aluminium is ……………. g/cc.
(a) 1.8 (b) 7.6 (c) 2.7 (d) 1

73. The m.pt of aluminium is ……………..
(a) 660°C (b) 660K (c) 1523°C (d) 1523 K

74. Aluminium is a ……………..
(a) bluish white (b) grey (c) silvery white (d) yellow

75. Aluminium is a self protecting metal due to the formation of its ……………….. on the surface.
(a) thick layer (b) oxide (c) chloride (d) carbonate

76. Aluminium decomposes ……………….. liberating hydrogen.
(a) cold water (b) hot water (c) boiling water (d) steam

77. Aluminium reacts with …………….. to liberate sulphur dioxide.
(a) hydrogen sulphide (b) dil. Sulphuric acid (c) sulphurous acid (d) hot con. Sulphuric acid

78. Aluminium was first isolated by …………….
(a) Baeyer (b) Charles Martin (c) Hall (d) Wohler

79. Aluminium was obtained by the electrolysis of ……………… by Charles Martin Hall in 1886.
(a) Bauxite (b) Cryolite (c) Corundum (d) Alumina

80. The chemical formula of bauxite is ………………
(a) Al_2O_3 (b) $Al_2O_3.2H_2O$ (c) Na_3AlF_6 (d) Al_2O_3

81. Which of the following is not an ore of aluminium ……………..
(a) Bauxite (b) cryolite (c) corundum (d) haematite

82. The principal ore of aluminium is …………..
(a) Bauxite (b) cryolite (c) corundum (d) haematite

83. Bauxite is purified by ……………….. process.
(a) Hall's (b) Baeyer's (c) Bessemer (d) Electrolytic

84. In Baeyer's process ………………. is added to act as a seeding agent.
(a) $Al(OH)_3$ (b) Al_2O_3 (c) NaOH (d) $NaAlO_2$

85. When sodium aluminate is agitated with water ………….. is precipitated.
(a) NaOH (b) Al_2O_3 (c) $Al(OH)_3$ (d) Fe_2O_3

86. When aluminium hydroxide is ignited …………….. is obtained.

(a) Al (b) Al_2O_3 (c) O_2 (d) H_2

87. Alumina is reduced to aluminium by …………….. process.

(a) Bayer's (b) Hall's (c) Bessemer (d) Castner's

88. The molten electrolyte in Hall's process is covered with a layer of …………….

(a) chalk (b) alumina (c) graphite (d) iron

89. In Hall's process, iron tank is lined with ………………

(a) alumina (b) graphite (c) cryolite (d) mercury

90. ………….. acts as anode in Hall's process.

(a) Iron (b) Graphite rods (c) Mercury (d) Aluminium

91. Blast furnace is made of ………………..

(a) steel (b) bricks (c) copper (d) stainless steel

92. Blast furnace is lined inside with ……………..

(a) fire resistant steel (b) fire resistant bricks (c) acid resisting bricks (d) flint

93. Blast furnace has ……….. arrangement for the introduction of charge.

(a) hopper (b) tuyers (c) slaghole (d) cup and cone

94. The base of blast furnace has a/an ……….. through which molten iron can be withdrawn.

(a) tapping hole (b) slaghole (c) tuyers (d) outlet

95. The small pipes through which a blast of hot air is admitted in blast furnace are called …………

(a) inlets (b) tuyers (c) slaghole (d) chambers

96. In the extraction of iron, flux used is …………..

(a) roasted ore (b) coke (c) silica (d) limestone

97. Density of iron is ………………g/cm^3.

(a) 17.8 (b) 11.2 (c) 10.3 (d) 7.86

98. Melting point of iron is …………..

(a) 1539 K (b) 1539^oC (c) 2085^oC (d) 2085K

99. Red hot iron reacts with ………….. to produce hydrogen.

(a) cold water (b) hot water (c) boiling water (d) steam

100. The percentage of carbon in cast iron is ……….

(a) 0.25% (b) 2-4.5% (c) 0.25 to 2% (d) zero percent

Answers:

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

II. Answer in one or two sentences:

1. Write the important ores of copper.

The important ores of copper are

1. Copper pyrites - $CuFeS_2$
2. Copper glance - Cu_2S
3. Cuprite - Cu_2O
4. Malachite - $Cu(OH)_2 \cdot CuCO_3$
5. Azurite - $2CuCO_3 \cdot Cu(OH)_2$

2. Give the important ores of Aluminium.

1. Bauxite ($Al_2O_3.2H_2O$) 2. Cryolite (Na_3AlF_6) 3. Corundum (Al_2O_3)

3. Name the steps involved in extraction of Aluminium from its ore.

The extraction of aluminium requires two stages:

(a) Purification of bauxite (b) Electrolytic reduction

4. Write a note on physical properties of Aluminium.

1. Aluminium is a silvery white metal.
2. It is extremely light and has a density 2.7g/cm^3.
3. It has a high tensile strength. It is malleable and ductile.
4. It melts at 660°C.
5. It is a good conductor of heat and electricity.

5. What happens when aluminium is treated with NaOH & KOH? Give equation.

Aluminium forms aluminates and liberates hydrogen on reaction with boiling aqueous NaOH and KOH.

$2Al + 2NaOH + 2H_2O \quad\quad 2NaAlO_2 + 3H_2\uparrow$

Sodium meta aluminate

$2Al + 2KOH + 2H_2O \quad\quad 2KAlO_2 + 3H_2\uparrow$

Potassium meta aluminate

6. Aluminium acts as a reducing agent – Explain with equation.

At high temperatures aluminium reduces chromic oxide to chromium (Alumino thermic process).

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

7. Give the composition of Alnico and give its uses.

Aluminium, iron, nickel, cobal.

Uses: Strong permanent magnets.

8. List the physical properties of copper?

1. It is a reddish brown metal.
2. It melts at 1083°C
3. It has a density 8.94g /cm^3
4. The metal is highly malleable and ductile.
5. It is a very good conductor of heat and electricity.

9. List the physical properties of iron.

1. Pure iron is grey-white metal and appears brown on rusting
2. It is very malleable and ductile.
3. It is good conductor of heat and electricity.
4. It has high density of 7.86g/cm^3 and a high melting point of 1539°C

10. Give any two uses of pig iron?

1. Pig iron is used for preparing objects like pipes, stoves and hot water radiators.
2. Wrought iron is used to make springs, tubes, locomotives, railway lines, electromagnets.

11. Name the different types of iron?

1. Cast iron, 2. Wrought iron and 3. Steel.

12. Define corrosion.

Corrosion may be defined as the slow and steady destruction of a metal or alloy by the environment.

It may be due to the chemical or electrochemical reaction of the metals with the substance present in the environment e.g. rusting of iron.

13. What is called rusting?

In case of iron, corrosion is called rusting. Rust is a hydrated ferric oxide represented as $Fe_2O_3.H_2O$.

14. Give the various causes of rusting?

Rusting of iron depends upon the following conditions:

(i) Impurities in iron *(ii)* Electrolytes *(iii)* Air and moisture

(iv) Acid forming gases such as CO_2 and SO_2 lead to rapid rusting during rainy day.

15. What is called anodizing?

In this process metals like aluminium, copper, *etc.,* are coated electrically with a thin and strong film of their oxides which protects them from rusting. Articles such as soap cases, handles, door knobs, *etc.,* are commonly anodized aluminium articles.

16. Write a note on electroplating?

It is a process of depositing stable metal (gold, silver) over a base metal (copper, iron). Iron can be coated with copper by electro deposition from a solution of copper sulphate. Spoons, handles of bicycles, taps and many other articles are electroplated. Many gift articles, medals, etc., are gold plated or silver plated electrolytically.

17. Give the composition of alnico & duralumin.

The composition of almino is Aluminium, iron, nickel and cobalt. The composition of duralumin is aluminium, copper, manganese and magnesium.

18. Give the uses of magnalium.

Magnalium is used in machine parts, balances, scientific instruments, etc.

19. What is ammonal?

A mixture of aluminium powder and ammonium nitrate is called ammonal and is used in explosives.

20. Give the equation for aluminothermic process.

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

21. What is used as an electrolyte in Hall's process?

Pure alumina dissolved in molten cryolite and mixed with a little amount of flourspar which lowers the melting point of the electrolyte is used.

22.Give equation for the reaction between iron and moist air.

In moist air, iron is oxidised to form rust

$$4Fe + 3O_2 + 3H_2O \longrightarrow \underbrace{Fe_2\,O_3 + 2Fe\,(OH)}_{\text{rust}}\,2$$

23. Name the various zones of blast furnace?

The various zones of blast furnace are (i) combustion zone, (ii) fusion zone, (iii) slag formation zone, (iv) reduction zone.

24. What is called calcination of iron ore?

The concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. During roasting (a) moisture is driven out and (b) impurities like sulphur, phosphorus, arsenic, etc, are oxidised off.

25. How is haematite concentrated?

The powdered ore is washed with a stream of water therefore the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

26. Give the important ores of iron?

The important ores of iron are

1. Haematite – Fe_2O_3 2. Siderite – $FeCO_3$
3. Magnetite – Fe_3O_4 4. Limonite – $2Fe_2O_3.\ 3H_2O$

27. What happens when copper is treated with dilute HNO_3 & concentrated HNO_3? Give equations.

With dilute nitric acid it gives **nitric oxide** whereas with concentrated nitric acid it gives mainly **nitrogen dioxide**.

$$\underset{\text{(dilute)}}{3Cu + 8\ HNO3} \longrightarrow 3\ Cu\ (NO_3)_2 + 2NO + 4H_2O$$

$$Cu + \underset{\text{(con.)}}{4\ HNO_3} \longrightarrow Cu\ (NO_3)_2 + 2NO_2 + 4H_2O$$

28. What is bessemerisation?

The process in which copper oxide reacts **with copper sulphide** to form copper metal is called bessemerisation.

29. How is copper pyrite roasted?

The concentrated copper pyrites ore is roasted in air in a blast furnace. On roasting, a part of copper sulphide (CuS) present in copper pyrites is oxidised to copper oxide (CuO) and sulphur dioxide gas is evolved.

$$2CuS + 3O_2 \longrightarrow 2CuO + 2SO_2$$

30. List the conditions required for rusting?

1 Rusting of iron depends upon the following conditions.

(i) Impurities in iron, (ii) Electrolytes. (iii) Air and moisture, (iv) Acid forming gases such as CO_2 and SO_2 lead to rapid rusting during rainy day.

31. Define corrosion?

Corrosion may be defined as the slow and steady destruction of a metal or alloy by the environment.

32. Name the different types of steel?

Different types of steel are (i) mild steel (ii) medium steel (iii) hard steel.

III. Answer in brief:

1. Discuss the various types of steel in detail.

Based on carbon content, there are three types of steel.

(i) Mild steel: It has the least carbon content; 0.1 to 0.15%. It is used for making wires and sheets.

(ii) Medium steel: It contains 0.2 to 0.5% carbon. It is harder than mild steel and is used for constructing rails, wheels etc.

(iii) Hard steel: This type of steel contains 0.5 to 1.5% carbon. It is very hard and used for making machine parts.

2. Explain tinning process.

This process involves the coating of tin (with molten tin) on cooking vessels made of copper and brass. These vessels get a greenish coating due to moist air or due to foodstuffs. **The greenish coating (due to corrosion) is poisonous**. Therefore, vessels are coated with tin. In case of tin coating on iron, the film will be effective as long as it is unscratched. When scratches occur at the coating surface both the metals are exposed to oxygen and rusting of iron occurs.

3. Give the action of copper with dilute and concentrated acids with equations.

Copper metal does not react with dilute HCl and H_2SO_4, However,

1. Copper reacts with **hot concentrated** sulphuric acid to form copper (II) sulphate and SO_2.

$$Cu + 2H_2SO_4 \longrightarrow \underset{\text{Copper (II) sulphate}}{CuSO_4} + 2H_2O + SO_2$$

2. With **dilute nitric acid** it gives nitric oxide where as with **concentrated nitric acid** it gives mainly nitrogen dioxide

$$3Cu + 8HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO^{-} + 4H_2O$$

(dilute)

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

(conc.)

4. Give the percentage composition and uses of 1. Stainless steel 2. Tungsten steel.

1. Stainless steel	11.5% Cr; 2% Ni	Resists corrosion	Common articles.
2. Tungsten steel	14-20% W; 3-8% Cr	Very hard	High speed tools.

5. Give the percentage composition and uses of 1. Alnico 2. Aluminium bronze.

1. Alnico	Aluminium, iron, nickel, cobalt	Strong permanent magnets
2. Aluminium bronze	Copper, aluminium	Cheap jewellery, cooking, vessel,

IV. Answer in detail:

1. What is called alloy? Tabulate the composition and uses of any three alloys Al and Fe.

An alloy is a homogeneous mixture of two or more metals or a metal and non-metals.

Aluminium combines with certain metals and form the following alloys.

Name of the alloys of Al	Composition	Uses
1. Duralumin	Aluminium, copper, manganese, magnesium	Air-craft parts, cars, ships, rails, space satellites
2. Magnalium	Aluminium and magnesium	Machine parts, balances, scientific instruments
3. Y –alloy	Aluminium, copper, nickel magnesium, silicon, iron	Lighter industrial parts

2. What are called special steels? Write the composition and the uses of the following special steels.

1. Invar 2. Chrome vanadium steel 3.Manganese steel

Steel mixed with small amount of nickel, cobalt, chromium, tungsten, molybdenum, manganese, etc., acquires special properties. Such products are called special steels or alloy steels. Some important alloy steels are described.

Name of the alloy steel	Metal added	Properties	Uses
1. Invar	36% Ni	Coefficient of expansion is very small	Measuring tapes, pendulums
2. Chrome – vanadium steel	1% Cr; 0.15% V	High tensile strength	Springs, shafts, axles.
3. Manganese steel	12-15% Mn	Hard and tough	Rock crushing machinery, safes, and helmets.

7 – NON -METALS

INTRODUCTION:

Elements are divided into two main classes: Metals and non-metals. This division is based upon differences in the characteristic properties of the two types of elements. Metals, in general, are solids (except mercury which is a liquid) having a bright appearance (metallic lustre) and are good conductors of heat and electricity. They are malleable (i.e. they can be flattened into thin sheets by hammering) and ductile (i.e. they can be drawn out into wire by stretching). On the other hand, some non-metals are gases, while others are solids (except bromine which is a liquid) which do not posses metallic lustre (except iodine). They are not malleable or ductile. In a chemical reaction, the metals have tendency to donate electrons. They posses low ionisation potentials and low electro negativities. Non-metals have tendency to accept or share electrons. They have high ionisation potentials and high electro negativities.

Non-metals are placed in the upper right hand corner of the periodic table (except hydrogen).

IMPORTANT TERMS & DEFINITIONS:

Non-Metals: Non –metals are the elements which form negative ions by gaining electrons. Hydrogen, oxygen, carbon, Sulphur, silicon and phosphorus are some of the common non –metals.

Inert Gases: Helium, neon, argon, krypton, xenon and radon are the inert gases which are also non – metals.

Position in the periodic table: The most metallic elements are on the extreme left side of the periodic table whereas non –metallic elements are on the extreme right side in the periodic table.

Electronic configuration: Non –metals have usually 4 to 8 electrons in the outermost shells of their atoms.

Physical Properties of Non Metals

Nature: Non –metals are brittle and bas conductor of heat and electricity (except graphite).

Melting and Boiling points: Non – metals have low melting and boiling points except graphite and diamond.

Silicon: It is the second most abundant element occurring in the earth's crust; the first being oxygen.

Silica: The simplest compound of silicon and oxygen is silicon dioxide.

Types of silica: Sand, quartz and opal.

Ferrosilicon: It is an alloy of silicon. It is used in the manufacture of apparatus for redistilling nitric acid.

Silico Bronze: It is used in the manufacture of telegraph and telephone wires.

Sodium Silicate: It is used for preserving eggs and for making chemical garden.

Silicones: They are polymeric organo silicon compounds having C –Si and Si –O – Si bonds.

Position in the periodic table: Silicon is an element of group IV (14 in modern periodic table) and occurs below the carbon in the periodic table.

Phosphorus

Electronic configuration: The electronic configuration of phosphorus is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^3$.

Position in the periodic table: Phosphorus is placed in group V of the periodic table below nitrogen.

Allotropic forms of phosphorus: White P, red P, scarlet p, metallic black P, violet P.

Phosphorescence: The property of glowing of white phosphorous in the dark.

Allotropy: Allotropy is a phenomenon in which the same element can exist in more than one crystalline or structural modification with change in physical properties.

Sulphur

Electronic configuration: The electronic configuration of Sulphur is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^4$.

Position in the periodic table: Sulphur is placed in-group VI (16 in modern periodic table) below oxygen.

Allotropes of Sulphur: (i) Rhombic Sulphur, (ii) Monoclinic or Prismatic Sulphur, (iii) Plastic or Amorphous Sulphur.

SELF EVALUATION (T.B.Page 126 & 127)

I. Choose the correct answer.

1.Non metals have .... electrons present in the outer most shells of their atoms.

(a) 4 to 8 (b) 1 to 3

(c) 2 to 4 (d) 1 to 8

2.Hydrogen, non-metal has only .... electron.

(a) 2 (b) 3

(c) 1 (d) 4

3.................... is the non metal found in a liquid state.

(a) Fluorine (b) Sulphur

(c) Silicon (d) Bromine

4.................... alloy is used in the manufacture of telegraph wires.

(a) Ferro- silicon (b) Silico- bronze

(c) Silicon carbide (d) Silicon chips

5.Sodium silicate solution is also called as ....

(a) Blue glass (b) Water-glass

(c) Watch- glass (d) Crystal-glass

6.Sulphur is extracted from the earth crust by the .... process.

(a) Bessemer's (b) Electro- refining

(c) Frasch (d) Electro thermal

7.Phosphorus exists in nature ....

(a) in free state (b) as phosphates

(c) as phosphoric acid (d) as phosphorus Pentoxide

8.The ore used in the extraction of phosphorus by modern electro thermal process is ....

(a) Bone ash (b) Rock phosphate

(c) Phosphoric acid (d) none of these

9.The molecular formula for carbonic acid is ....

(a) H_2SO_4 (b) CO_3

(c) H_2CO_3 (d) HCOOH

10.An allotrope of phosphorus, which does not exhibit phosphorescence, is ....

(a) Red phosphorus (b) White phosphorus

(c) Both (d) None

Answers:

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

II. Answer the following in One or Two sentences.(T.B. Page 128)

1.Where are non-metals located in the periodic table?

The most metallic elements are on the extreme left side of the periodic table where as non-metallic elements are on the extreme right side in the periodic table. Only one non-metal hydrogen (H) has been placed on the left side in the periodic table.

2.Define allotropy.

Allotropy is a phenomenon in which the same element can exist in more than one crystalline or structural modification with change in physical properties.

3.Name different minerals of phosphorus.

The different minerals containing phosphorus are:

1.Phosphorite- $Ca_3(PO_4)_2$ 2.Chloropatite -$3Ca_3(PO_4)_2.CaCl_2$ 3.Apatite-$3Ca_3(PO_4)_2.CaF_2$

4. What happens when phosphorous react with alkalies? Give a balanced chemical equation?

White phosphorus reacts with hot solutions of caustic soda or caustic potash giving phosphine gas.

$$P_4 + 3NaOH + 3H_2O \longrightarrow \underset{\text{phosphine}}{PH_3\uparrow} + \underset{\text{sodium hypophosphite}}{3NaH_2PO_2}$$

5.What is the action of sulphur with sulphuric acid? Give equation.

Hot and concentrated sulphuric acid oxidises sulphur to sulphur dioxide

$$\underset{\text{Sulphur}}{S} + \underset{\text{Sulphuric acid (Hot and conc)}}{2H_2SO_4} \longrightarrow \underset{\text{Sulphur dioxide}}{3SO_2} + \underset{\text{Water}}{2H_2O}$$

6.What is the action of sulphur with nitric acid? Give equation.

Sulphur is oxidised to sulphuric acid by hot and concentrated nitric acid.

S	+	$6HNO_3$	$\longrightarrow$	H_2SO_4	+	$6NO_2$	+	$2H_2O$
Sulphur		Nitric acid (Hot and conc.)		Sulphuric acid		Nitrogen dioxide		

III. Answer in brief.(T.B. Page 128)

1.Give any five physical properties of non-metals.

1. Non metals are brittle and cannot be used to make sheets or wires.
2. Non metals are bad conductors of heat and electricity except graphite.
3. Most of the solid non-metals are quite soft except diamond.
4. Non metals can be easily broken, *i.e.,* the tensile strength of non metals is low.
5. Non metals have low melting and boiling points. The only exception is graphite, whose melting point is very high.

2.What is super phosphate? Give its preparation.

Super phosphate fertilizer is a phosphatic fertilizer which is soluble in water. Super phosphate is calcium dihydrogen phosphate ($Ca(H_2PO_4)_2$).

Super phosphate fertilizer is prepared by heating rock phosphate mineral (containing calcium phosphate) with a calculated quantity of concentrated sulphuric acid.

$Ca_3(PO_4)_2$	+	$2H_2SO_4$	$\longrightarrow$	$Ca(H_2PO_4)_2$	+	$2CaSO_4$
Calcium phosphate		Sulphuric Acid		calcium dihydrogen phosphate		calcium sulphate

3.How is silicon prepared from silicon dioxide?

1. **Commercial scale preparation:** Silicon is prepared by heating finely powered silicon dioxide (sand) with coke (carbon) in an electric furnace.

SiO_2	+	$2C$	$\longrightarrow$	Si	+	$2CO\uparrow$
Silicon dioxide		coke		Silicon		Carbon monoxide

4.Give the different uses phosphorus.

1. It is used in match industry.
2. It is used in manufacture of fertilizer like super phosphate etc.
3. It is used as rat poison.
4. It is used in fire-works, smoke bombs.
5. It is used in radioactive therapy.

5.Give the laboratory method of preparation of silicon.

Silicon is prepared in the laboratory by heating powdered silicon dioxide with magnesium powder.

SiO_2	+	$2Mg$	$\longrightarrow$	Si	+	$2MgO$
				Silicon		Magnesium oxide

6.Give any three uses of sulphur.

1. Sulphur is used in the manufacture of sulphuric acid (king of chemicals).
2. Sulphur is used in the manufacture of carbon disulphide- a solvent used in the manufacture of rayon.
3. Sulphur is used as an antiseptic in making skin ointments (skin creams).

IV. Answer in detail.(T.B. Page 128 & 129)

1.Explain the chemical properties of non-metals?

1. **Reaction with oxygen:** Nonmetals like carbon and sulphur combine with oxygen to form CO_2 and SO_2 respectively. These oxides dissolve in water to give carbonic acid and sulphurous acid.

$$C + O_2 \longrightarrow CO_2(g)$$

$$CO_2(g) + H_2O \longrightarrow H_2CO_3$$

(Carbonic acid)

$$S + O_2 \longrightarrow SO_2(g)$$

$$SO_2(g) + H_2O \longrightarrow H_2SO_3$$

(Sulphurous acid)

2. Reaction with acids: Sulphur reacts with concentrated nitric acid to form H_2SO_4 and NO_2.

$$S + 6HNO_3 \longrightarrow H_2SO_4 + 6NO_2(g) + 2H_2O$$

3. Reaction with chlorine: Nonmetals like phosphorus and hydrogen reacts with chlorine to form phosphorus trichloride and hydrogen chloride respectively.

$$P_4 + 6Cl_2 \longrightarrow 4PCl_3(g)$$

$$H_2 + Cl_2 \longrightarrow 2HCl\ (g)$$

4. Reaction with hydrogen: Nonmetals like sulphur, nitrogen reacts with hydrogen to give hydrogen sulphide and ammonia respectively.

$$H_2 + S \longrightarrow H_2S\ (g)$$

$$N_2 + 3H_2 \longrightarrow 2NH_3(g)$$

2.Explain the reaction of silicon with - (a) air; (b) chlorine; (c) steam; (d) Hydrochloric acid and (e) Sodium hydroxide.

(a) Reaction with oxygen: Silicon burns in the oxygen of air to form a white solid compound called silicon dioxide(silica/sand).

$$Si + O_2 \quad SiO_2$$

Silicon dioxide

(b) Reaction with chlorine: On heating to a temperature of 450°C, silicon combines with chlorine to form silicon tetra chloride.

$$Si + 2Cl_2 \quad \underset{450°C}{} \quad SiCl_4$$

Silicon tetra chloride

(c) Reaction with steam: When red hot silicon reacts with steam to form silicon dioxide and hydrogen gas.

$$\underset{\text{(Red hot)}}{\underset{\text{Silicon}}{Si}} + \underset{\text{Steam}}{2H_2O} \longrightarrow \underset{\text{dioxide}}{\underset{\text{Silicon}}{SiO_2}} + \underset{\text{Hydrogen}}{2H_2\uparrow}$$

(d) Reaction with Hydrochloric Acid: When silicon reacts with hydrochloric acid to form silicon tetra chloride and hydrogen gas is liberated.

$$Si + 4HCl \longrightarrow \underset{\text{tetrachloride}}{\underset{\text{Silicon}}{SiCl_4}} + \underset{\text{Hydrogen}}{2H_2\uparrow}$$

(e) Reaction with sodium hydroxide: Silicon reacts with hot sodium hydroxide solution to form sodium silicate with the evolution of hydrogen gas.

$$Si + 2NaOH + H_2O \longrightarrow \underset{\text{silicate}}{\underset{\text{Sodium}}{Na_2SiO_3}} + \underset{\text{Hydrogen}}{2H_2\uparrow}$$

3.What is allotropy? Explain the different allotrophic forms of sulphur?

Allotropy: Allotropy is a phenomenon in which the same element can exist in more than one crystalline or structural modification with change in physical properties.

Sulphur exists in three allotrophic forms. They are *(i)* Rhombic sulphur *(ii)* Monoclinic or Prismatic sulphur and *(iii)* Plastic sulphur or Amorphous sulphur. The first two are in crystalline form while the third is amorphous.

(i) Rhombic sulphur: This is obtained by dissolving the extracted roll sulphur in carbon disulphide and slowly allowed to evaporate in air slowly.

Rhombic sulphur crystallises in the shape like two pyramids joined at the base. It is a pale yellow crystalline solid. It is insoluble in water but soluble in carbon disulphide and benzene. It is stable below 96°C. Above 96°C rhombic sulphur changes into prismatic sulphur.

Fig. Rhombic sulphur

(ii) Monoclinic sulphur: It is prepared by melting rhombic sulphur in a dish and allowing it to cool until a crust is formed at the surface. Two holes are then pierced into the crust. The liquid sulphur lying below the crust (which has not yet solidified) is poured out through one of the holes. On removing the crust, small needle like crystals of monoclinic sulphur becomes visible. It melts at 120°C. It is soluble in carbon disulphide.

(iii) Plastic sulphur: It is prepared by boiling molten sulphur up to 445°C. When boiling sulphur is cooled suddenly by pouring into cold water, a brown rubber- like mass of plastic sulphur is obtained. It is plastic like in texture. It is insoluble both in water and carbon disulphide.

4. What are the ores of phosphorus? How is phosphorus extracted by the modern electric thermal process?

The major minerals containing phosphorus are:

1.Phosphorite-$Ca_3(PO_4)_2$ 2.Chloropatite-$3Ca_3(PO_4)_2.CaCl_2$ 3.Apatite-$3Ca_3(PO_4)_2.CaF_2$

Phosphorus is an essential constituent of bones and is also present in blood, brain and other parts of the body. It is also necessary for the growth of plants. As phosphoproteins, it is present in milk and eggs.

Rock phosphate is mainly calcium phosphate where as bone ash (animal charcoal) contains about 80% of calcium phosphate.

Extraction of Phosphorus: (Modern electro thermal process)

In this process phosphorous is extracted by heating bone ash or phosphate rock, phosphorite, [$Ca_3(PO_4)_2$], with sand (SiO_2) and coke in an electric furnace at about 1500°C.

A mixture of powdered phosphate rock, sand and coke (charge) is introduced into the furnace through the hooper by means of worm conveyor. The electric arc is set up between the electrodes E, E′ produce the high temperature required for the reaction at about 1500°C.

(i) Silica reacts with calcium phosphate to give phosphorus pentoxide is formed.

$$\underset{\text{Calcium Phosphate}}{Ca_3(PO_4)_2} + \underset{\text{Silica}}{3SiO_2} \longrightarrow \underset{\text{Calcium Silicate}}{3CaSiO_3} + \underset{\text{Phosphorus pentoxide}}{P_2O_5}$$

(ii) Phosphorus pentoxide in turn, is reduced by coke to give phosphorus vapour at about 1500°C.

$$2P_2O_5 + 10C \longrightarrow \underset{\text{phosphorus}}{P_4\uparrow} + \underset{\text{carbon monoxide}}{10CO\uparrow}$$

Phosphorus vapour escapes through the outlet and is condensed by cold water in condenser. Impure phosphorus obtained flows into the collector as shown.

Fig. Extraction of Phosphorus

Calcium silicate collects at the base of the furnace in the molten form as slag and is removed through the pipe S.

The impure solid phosphorus is purified by melting under a solution of chromic acid and allowed to remain there for few hours. The bottom layer of phosphorus is separated, washed with hot water and filtered through canvas cloth or chamois leather. Phosphorus is then cast into sticks by means of tube of glass or tin standing in cold water. The phosphorus sticks are kept under water to protect them from air oxidation.

5. What happens when phosphorus is (a) exposed to excess air; (b) exposed to chlorine; (c) heated with concentrated nitric acid and (d) Heated with caustic alkali solution.

(a) Exposed to excess air (Combustion): Phosphorus burns in air forming phosphorus trioxide and phosphrous pentoxide.

$$P_4 + 3O_2 \longrightarrow P_4O_6$$

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

(b) Exposed to chlorine (reaction with chlorine): White phosphorus spontaneously catches fire in chlorine forming phosphorus trichloride and pentachloride while the red phosphorus catches fire in chlorine only on heating

$$P_4 + 6Cl_2 \longrightarrow \underset{\text{Phosphorus trichloride}}{4PCl_3}$$

$$P_4 + 10Cl_2 \longrightarrow \underset{\text{Phosphorus pentachloride}}{4PCl_5}$$

(c) Heated with concentrated nitric acid (Reaction with acids): White phosphorus is heated with concentrated nitric acid, it is oxidized to orthophosphoric acid H_3PO_4.

Red phosphorus reacts in presence of iodine as catalyst.

$$P + 5HNO_3 \longrightarrow \underset{\text{Orthophosphoric Acid}}{H_3PO_4} + 5NO_2 + H_2O$$

(d) Heated with caustic alkali solution (Reaction with alkalis): White phosphorus reacts with hot solutions of caustic soda or caustic potash giving phosphine gas.

$$P_4 + 3NaOH + 3H_2O \longrightarrow \underset{\text{phosphine}}{PH_3\uparrow} + \underset{\text{sodium hypophosphite}}{3NaH_2PO_2}$$

6.How do you extract sulphur by Frasch process?

In the Frasch process, three concentric pipes are sunk to the sulphur beds, 500 to 800 feet below the surface of the earth. Water is heated to 180°C under pressure is forced through the outermost pipe. While hot, compressed air is forced down the inner most pipe. The hot water melts the sulphur and a foam of water, sulphur and air is produced. This foam is brought up to the surface through the central pipe by the action of air blast. The molten sulphur is run into separating tanks where it is separated from air and water. The sulphur so obtained is 99.5% pure.

Fig. Frasch process

7.Discuss the various allotropes of sulphur in detail.

Sulphur exists in three allotrophic forms. They are *(i)* Rhombic sulphur *(ii)* Monoclinic or Prismatic sulphur and *(iii)* Plastic sulphur or Amorphous sulphur.

(i) Rhombic sulphur: This is obtained by dissolving the extracted roll sulphur in carbon disulphide and slowly allowed to evaporate in air slowly.

Rhombic sulphur crystallises in the shape like two pyramids joined at the base. It is a pale yellow crystalline solid. It is insoluble in water but soluble in carbon disulphide and benzene. It is stable below 96°C. Above 96°C rhombic sulphur changes into prismatic sulphur.

Fig. Rhombic sulphur

(ii) Monoclinic sulphur: It is prepared by melting rhombic sulphur in a dish and allowing it to cool until a crust is formed at the surface. Two holes are then pierced into the crust. The liquid sulphur lying below the crust (which has not yet solidified) is poured out through one of the holes. On removing the crust, small needle like crystals of monoclinic sulphur becomes visible. It melts at 120°C. It is soluble in carbon disulphide.

(iii) Plastic sulphur: It is prepared by boiling molten sulphur up to 445°C. When boiling sulphur is cooled suddenly by pouring into cold water, a brown rubber- like mass of plastic sulphur is obtained. It is plastic like in texture. It is insoluble both in water and carbon disulphide.

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

7. NON – METAL

OBJECTIVE TYPE QUESTIONS

1. Phosphorus is manufactured by heating by heating in a furnace
   (a) Bone ash, sodium chloride and coke (b) Bone ash, silica and coke
   (c) Bone ash, silica and lime (d) Bone ash, coke and lime stone.
2. Super phosphate of lime contains
   (a) $Ca_3 (PO_4)_2$ (b) $CaHPO_4$ (c) $Ca_3 (PO_4)_2 + H_3 PO_4$ (d) $Ca(H_2PO_4)_2$.
3. The basicity of H_3PO_4 is
   (a) 2 (b0 3 (c) 4 (d) 5.
4. Phosphine is generally prepared in the laboratory
   (a) By heating phosphorous in a current of hydrogen
   (b) By decomposition of P_2H_4 at 110°C
   (c) By heating red phosphorous with an aqueous solution of caustic soda
   (d) By heating white phosphorous with caustic potash.
5. White phosphorus is generally preserved in
   (a) alcohol (b) water (c) kerosene oil (d) ether.
6. White P when boiled with strong solution of caustic soda produces
   (a) Phosphine (b) Posh acid (c) Phosphorous acid (d) None.
7. Phosphine is prepared by the action of
   (a) P and H_2SO_4 (b) P and Na OH (c) P and H_2S (d) P and HNO_3.
8. Which of the P is most stable?
   (a) Red (b) White (c) Scarlet (d) All stable.
9. In modern process, white phosphorus is manufactured by
   (a) Heating a mixture of phosphorite mineral with sand and coke in an electric furnace
   (b) Heating calcium phosphate with lime
   (c) Heating bone ash with coke

(d) Heating phosphate mineral with sand.

10. Which of the following phosphorus is most reactive?
(a) Red phosphorus (b) white phosphorus (c) Scarlet phosphorus (d) Violet phosphorus

11. White phosphorus is
(a) Brittle power (b) Less reactive (c) Very poisonous (d) A linear diatomic molecule.

12. The molecular formula of sulphur is
(a) S_2 (b) S_4 (c) S_6 (d) S_8

13. Which of the following process is used for extraction of sulphur from sulphur beds?
(a) Acheson process (b) Carter process (c) Frasch process (d) Le - Blanc process.

14. In the Frasch process, molten sulphur rises up from the
(a) Inner pipe (b) Outer pipe (c) Middle pipe (d) All of these.

15. Rhombic sulphur on heating at above 96°
(a) Sublimes.
(b) Melts to a thin pale yellow liquid and then vapourises.
(c) Melts to a thin dark coloured liquid and then vapourises.
(d) Changes into prismatic sulphur.

16. Which of the following elements occur in abundantly?
(a) Si (b) Ge (c) sn (d) C.

17. SiO_2 or ordinary sand reacts with
(a) Conc. HC*l* (b) HF (c) NaOH(hot) (d) C.

18. The atomicity of sulphur is ………….
(a) 4 (b) 6 (c) 8 (d) one

19. …………… is used as a disinfectant.
(a) C (b) S (c) SO_2 (d) SO_3

20. …………….. is used in vulcanisation
(a) S (b) rubber (c) carbon (d) H_2SO_4

21. …………… is used as an antiseptic in making skin oinments.
(a) S (b)CS_2 (c) H_2SO_4 (d)HCl

22. ……………… is used as solvent in the manufacture of rayon.
(a) S (b) CS_2 (c) H_2SO_4 (d) HNO_3

23. The king of chemicals is ……………

(a) H_2O (b) HNO_3 (c) H_2SO_4 (d) HCl + HNO_3

24. Sulphur is oxidised to …………….by hot and con. nitric acid.

(a) SO_2 (b) SO_3 (c) H_2SO_3 (d) H_2SO_4

25. Sulphur reacts with Con. H_2SO_4 toproduce……….

(a) SO_2 (b) SO_3 (c) H_2S (d) H_2SO_3

26. At a temperature of about………….. sulphur becomes highly viscous.

(a) 160°C (b) 120°C (c) 220°C (d) 444°C

27.The slag in the manufacture of phosphorus is ………….

(a) $CaSiO_3$ (b) $CaCO_3$ (c) coke (d) $Ca_3(PO_4)_2$

28. Phosphorus is filtered through……………..

(a) canvas (b) chamois leather (c) filter paper (d) both (a) and (b)

29. White phosphorus is kept under………..

(a) water (b) kerosene (c) benzene (d) alcohol

30. Which one of the following does not exist in allotropic form……….

(a) P (b) C (c) S (d) Na

31. The most reactiv form of phosphorus is……………

(a) White P (b) red P (c) scarlet P (d) violet P

32. White phosphorus has ………………….smell.

(a) pungent (b) garlic (c) fishy (d) rotten egg

33. The density of red phosphorus is……………

(a) $1.8 g/cm^3$ (b) $2.3 g/cm^3$ (c) $7.6 g/cm^3$ (d) $1 g/cm^3$

34. Phosphorus is placed next to …………. In V group.

(a) carbon (b) oxygen (c) sulphur (d) nitrogen

35. The electronic configuration of phosphorus is………………..

(a) $1s^2, 2s^2, 2s^3$ (b) $1s^2, 2s^2, 2p^6, 3s^2$ (c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ (d) $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

36. Bone ash contains about ………….. of calcium phosphate.

(a) 50% (b) 60% (c) 75% (d) 80%

37. In modern electro thermal process, a mixture of powdered phosphate rock, sand and coke is introduced into the furnace through the hooper by means of ………..

(a) cup and cone arrangement (b) worm conveyor

(c) funnel (d) tuyers

38. The temperature produced in modern electric furnace is about………..

(a) 875°C (b) 1200°C (c) 1500°C (d) above 2000°C

39. Silica reacts with calcium phosphate to produce ………….

(a) P_2O_3 (b) P (c) P_4 (d) P_2O_5

40. Coke reacts with P_2O_5 to produce………..

(a) CO_2 (b) P_2O_3 (c) CO (d) H_3PO_3

41.Silicon combines with chlorine at a temperature of…………..

(a)450 K (b) 450°C (c) 1020 K (d) 1020 °C

42. …………… is rather inert at room temperature.

(a) Carbon (b) Sulphur (c) Silicon (d) Boron

43. The reducing agent used to reduce SiO_2 to Si is………..

(a) C (b) Mg (c) K (d) both (a) and (b)

44. Silicon is present in group four, next to………..

(a) nitrogen (b) carbon (c) oxygen (d) fluorine

II. ANSWER THE FOLLOWING IN ONE OR TWO SENTENCES.

1. Give the reaction between Silicon with Hydrochloric acid.

When silicon reacts with hydrochloric acid to form silicon tetra chloride and hydrogen gas is liberated.

$$\underset{}{Si} + 4HCl \longrightarrow \underset{\text{Silicon tetrachloride}}{SiCl_4} + \underset{\text{Hydrogen}}{2H_2 \uparrow}$$

2. What happens when phosphorus is heated with concentrated nitric acid?

When white phosphorus is heated with concentrated nitric acid, it is oxidized to orthophosphoric acid (H_3PO_4.) Red phosphorus reacts in presence of iodine as catalyst.

$$P + 5HNO_3 \longrightarrow \underset{\text{Orthophosphoric Acid}}{H_3PO_4} + 5NO_2 + H_2O$$

3. Name the various allotropes of sulphur.

Sulphur exists in three allotrophic forms. They are *(a)* Rhombic sulphur *(b)* Monoclinic or Prismatic sulphur and *(c)* Plastic sulphur or Amorphous sulphur. The first two are in crystalline form while the third is amorphous.

4. Draw the structure of sulphur molecules.

Fig. Sulphur molecule

5. What happens when sulphur reacts with Carbon and hydrogen?

$$\underset{\text{Carbon}}{C} + \underset{\text{Sulphur}}{2S} \longrightarrow \underset{\text{Carbon disulphide}}{CS_2}$$

$$\underset{\text{Hydrogen}}{H_2} + \underset{\text{Sulphur}}{S} \longrightarrow \underset{\text{Hydrogen sulphide}}{H_2S}$$

6. How many non-metals are known today? Name the non-metal which is placed on the left side in the periodic table?

There are twenty two non-metals known. Hydrogen is placed on the left side in the periodic table.

7. How many non-metals are present in solid state & gaseous state?

There are 10 non-metals in solid state and 11 non-metals in gaseous state.

8. Give example for a liquid non-metal & a non metal which conduct electricity.

The liquid non-metal is bromine. Graphite is a good conductor of electricity.

9. Name the different forms in which silica exists?

Silica exists in a variety of forms such as sand, quartz and opal.

10. What is called waterglass? Give its use.

Sodium silicate is called water glass. It is used for preserving eggs and for making chemical gardens or silica gardens.

11. What is the atomicity of sulphur?

One molecule of sulphur has 8 sulphur atoms at ordinary temperature, so it is written as S_8. The atomicity of sulphur is 8.

12. What is Vulcanisation?

The process of heating crude rubber with sulphur to 140^0C is called Vulcanisation.

13. Give any three allotropes of sulphur.

The three allotropes of sulphur are (a) Rhombic sulphur (b) Monoclinic sulphur and (c) Plastic or Amorphous sulphur.

14. Give the position of sulphur in the periodic table?

Sulphur is present in the periodic table in the periodic table in group VI along the period 3.

15. Give the position of phosphorus in the periodic table?

Phosphorus is placed in group 5 along the third period in the periodic table.

16. Give the various allotropic modifications of phosphorus?

The various allotropic modifications of phosphorus are (1) White phosphorus (2) Red phosphorus (3) Scarlet phosphorus (4) Metallic black phosphorus (5) Violet phosphorus.

17. What is called phosphorescence?

The property of glowing of white phosphorus in the dark is known as phosphorescence.

18. What are silicones? Give the nature and general formula.

Silicones are polymeric organosilicon compounds. Silicones are chemically inert and thermally stable compounds. They have the general formula $(R_2SiO)_n$.

19. Give the uses of silicon.

Silicones are used in cosmetics, as lubricants as water repellent, as antifoams, as stop-cock grease, as polish, as insulating material for electrical motors and other electrical appliances.

III. Answer in brief:

1. List the various allotropes of phosphorus.

The various allotropic modification of phosphorus are:

1. White phosphorus
   (a) α-White phosphorus
   (b) β -White phosphorus
2. Red phosphorus
3. Scarlet phosphorus
4. Metallic black phosphorus
5. Violet phosphorus.
   (a) α - Black phosphorus
   (b) β - Black phosphorus

2. Write a brief note on silicon polymer.

Silicones are polymeric organosilicon compounds having C – Si and Si – O – Si bonds. They have the general formula $(R_2SiO)_n$ (Where R = methyl or phenyl). The general structure of silicone polymer is represented as follows:

```
        R          R          R          R
        |          |          |          |
   —O—Si—    O—Si—    O—Si—    O—Si—O—
        |          |          |          |
        R          R          R          R
```

(Silicon Polymer) (R = Methyl or Phenyl)

Silicones are chemically inert and thermally stable compounds. They are strongly water repelling substances which can be converted into oils, resins and rubbery elastomers. Silicones are used in cosmetics, as lubricants, as water repellent, as antifoams, as stop-cock grease, as polish, as insulating material for electrical motors and other electrical appliances.

3. Give any three uses of Silicon.

1. Silicon is used in the preparation of some important alloys e.g. Ferro – silicon, silico bronze.
   Ferro- silicon, is used in the manufactured of apparatus for redistilling nitric acid.
   Silico-bronze is used in the manufacture of telegraph and telephone wires.
2. Silicon dioxide is used in the making glass and cement quartz crystals are used in matches.
3. Silicon is used in making micro – processors or "silicon chips".

4. Discuss the position of non-metals in the periodic table?

The most metallic elements are on the extreme left side of the periodic table whereas non-metallic elements are on the extreme right side in the periodic table. Only one non-metal hydrogen (H) has been placed on the left side in the periodic table.

5. Give the general electronic configuration of non-metals?

Non-metals have usually 4 to 8 electrons in the outermost shells of their atoms. For e.g., carbon, nitrogen, oxygen, fluorine and neon have 4,5,6,7 and 8 electrons in their outermost orbit.

6. Write the reaction between a non-metal and oxygen. Give the nature of the oxides formed.

Non-metals like carbon and sulphur combine with oxygen to form CO_2 and SO_2 respectively. These oxides dissolve in water to give carbonic acid and sulphurous acid. Hence they are acidic in nature.

$$C + O_2 \longrightarrow CO_2\ (g)$$
$$CO_2\ (g) + H_2O \longrightarrow H_2CO_3$$

$$S + O_2 \longrightarrow SO_2\ (g)$$
$$SO_2 + H_2O \longrightarrow H_2SO_3$$

IV. Answer in detail:

1. Discuss the action of heat on sulphur at various temperatures.

1. Sulphur melts at 115°C to form a pale yellow mobile liquid.
2. On heating further to about 160°C instead of becoming more mobile, the yellow liquid sulphur becomes dark brown and highly viscous (very thick).
3. On further heating, the liquid sulphur becomes lighter in colour and less viscous.
4. Finally, the liquid sulphur boils at 444°C giving away yellow brown vapours. These changes are reversed on cooling.

The unusual behaviour of sulphur i.e. the change in the viscosity of the liquid is due to the opening of the S_8 rings into chains.

2. Tabulate the difference between white and red phosphorus.

	Property	White Phosphorus	Red Phosphorus
1.	State and colour	White solid when freshly cut changes to yellow on keeping.	Opaque red solid
2.	Odour	Like garlic.	Odourless
3.	Hardness	Soft, can be cut with a knife	Brittle powder
4.	Density	1.8g / cm³ (lighter)	2.3g /cm³ (Heavier)
5.	Ignition temperature	30°C (low)	260°C (high)
6.	Melting point	44°C (under water)	Does not melt, sublimes at 400°C
7.	Solubility	Soluble in carbon disulphide	Insoluble in carbon disulphide
8.	Physiological action	Very poisonous	Non- poisonous
9.	Chemical reactivity	Very reactive	Less reactive
10.	Action of air	Burns spontaneously in air, show's phosphorescence.	Stable in air, does not show phosphorescence.
11.	Action of chlorine	Combines spontaneously with chlorine to give PCl_3 and PCl_5	Combines with chlorine only on heating to give PCl_3 and PCl_5
12.	Action of hot sodium hydroxide solution.	Phosphine gas is evolved	No action

8 – ALCOHOLS

INTRODUCTION

When one or more hydrogen atoms in an aliphatic hydrocarbon are replaced by one or more hydroxyl (OH) groups, the resulting compounds are known as alcohols.

$$\underset{\text{Alkane}}{R-H} \xrightarrow[-H]{+OH} \underset{\text{Alcohol}}{R-OH}$$

According to the number of the hydroxyl groups present in them, they are known as ***monohydric, dihydric, trihydric*** **or** ***polyhydric*** alcohols. In case of polyhydric alcohols (having more than one OH group), the hydroxyl groups are always present on different carbon atoms. It is because the alcohols having two or more OH groups attached to the same carbon are unstable.

IMPORTANT TERMS & DEFINITIONS:

Alcohols: Alcohols are organic compounds with general formula R – OH where R is an alkyl group.
Hydroxy derivatives: Alcohols may be considered as hydroxy derivatives of hydrocarbons in which one or more hydrogen atoms are replaced by hydroxyl (-OH) group.
Aliphatic alcohols: The open chain alcohols are called aliphatic alcohols where R is an alkyl group.
Classification: Alcohols are classified as mono, di, tri and polyhydric alcohols.
Nomenclature: There are two methods for naming alcohols. (a) Common Name (b) IUPAC name
Fermentation: It is a slow anaerobic decomposition of big organic molecules into simpler ones under the catalytic influence of non-living complex substances called ferments. In many causes enzymes acts as ferments.
Molasses: It is the mother liquor left over after the crystallization of sugar from the sugarcane juice. It is a dark colored syrupy liquid containing 50 – 55 % total sugars (such as sucrose, glucose and fructose).
Wash: The filtrate collected in fermentation contains almost 10% of ethyl alcohol, called wash.
Azeotropic mixture: It is a mixture, which boils at a constant temperature and distills over completely at the same temperature without change in composition.
Denatured alcohol: Rectified spirit is made unfit for drinking purpose, by adding 5% methyl alcohol (poison), 0.5% pyridine (unpleasant odour) and some colouring matter (methyl violet dye). It is called denatured alcohol or **methylated spirit**.
Power alcohol: Rectified spirit does not mix properly with petrol. Hence, it is mixed with ether or benzene. One part of this mixture is added to four parts of petrol. This is called power alcohol or gasohol.
Esterification: The process of the formation of an ester (ethyl ethanoate) by the combination of ethanol with ethanoic acid is known as **esterification.**
Detection of Alcohol: Alcohol can be detected by the following tests.
Sodium metal test: when a small piece of sodium is added to an alcohol, hydrogen gas is evolved with effervescence.
Phosphorous pentachloride test: On treatment with phosphorous pentachloride, alcohols become warm and hydrogen chloride gas is evolved.

SELF EVAULATION (T.B.Page 137 & 138)

I. Choose the correct answer.

1. The enzyme used to convert Glucose to C_2H_5 OH is ....
   (a) Zymase (b) Invertase
   (c) Both (a) & (b) (d) None
2. Alcohols is a ....
   (a) Non- conductor (b) conductor
   (c) Semiconductor (d) Insulator
3. Denatured alcohol contains .... methanol.
   (a) 5% (b) 0.5%
   (c) 25% (d) 15%
4. The general formula for alcohol is ....
   (a) R – O – R (b) R – OH
   (c) R – COOH (d) R – CHO

5. An example for secondary alcohol is ....
   (a) Propan -1-ol (b) Propan-2-ol
   (c) Butan -1-ol (d) Butan-2-ol
6. Ethanol is used as an ....
   (a) Antiseptic (b) antipyretic
   (c) anaesthetic (d) None
7. Rectified spirit consists of .... of alcohol.
   (a) 95.5% (b) 95%
   (c) 98.5% (d) 98%
8. 2-methyl propan-2-ol is an example for ....
   (a) Primary alcohol (b) secondary alcohol
   (c) Tertiary alcohol (d) None
9. .... is added as a food for yeast.
   (a) Ammonium sulphate (b) ammonium phosphate
   (c) both (a) or (b) (d) None
10. Butan-2-ol is a ....
   (a) Primary alcohol (b) Secondary alcohol
   (c) Tertiary alcohol (d) aldehyde

Answers:
1. (a) 2. (a) 3. (a) 4. (b) 5. (d) 6. (a) 7. (a) 8. (c) 9. (c) 10. (b)

II. Answer the following in One or Two sentences.(T.B.Page 138)

1.Define Molasses.

Molasses is the mother liquor left over after the crystallization of sugar from the sugarcane juice.

2.Mention the role of H_2SO_4 in the manufacture of ethanol.

Dilute sulphuric acid is added to check the bacterial growth and to bring pH to 4 – 5 (Yeast thrives in acidic solution) and to stop the growth of unwanted bacteria.

3.What is called absolute alcohol?

Rectified spirit is mixed with quick lime (CaO) and allowed to stand for some time. It is then distilled. This is called lime of alcohol (98%). Final traces of water are removed by distillation with calcium metal to get absolute alcohol (100% alcohol). Anhydrous copper sulphate can also be used to remove final traces of water.

4.What is called power alcohol?

Rectified spirit does not mix properly with petrol. Hence, it is mixed with ether or benzene. One part of this mixture is added to four parts of petrol. This is called power alcohol.

5.What is called denatured alcohol?

Rectified spirit is made unfit for drinking purpose, by adding 5% methyl alcohol (poison), 0.5% pyridine (unpleasant odour) and some colouring matter (methyl violet dye). It is called denatured alcohol or **methylated spirit**. It is used for the preparation of paints and varnishes as a solvent.

6.Define azeotropic mixture.

Azeotropic mixture is a mixture, which boils at a constant temperature and distills over completely at the same temperature without change in composition.

7.$C_2H_5OH + PCl_5 \longrightarrow$ ....?......... +........?......... +........?......... Complete and balance the reaction.

$C_2H_5OH + PCl_5 \longrightarrow \mathbf{C_2H_5Cl + POCl_3 + HCl}$

III. Answer in brief. (T.B. Page 138)

1.Define fermentation with an example.

Fermentation is a slow anaerobic decomposition of big organic molecules into simpler ones under the catalytic influence of non-living complex substances called ferments. In many causes enzymes acts as ferments.

Example: Manufacture of ethyl alcohol from sugar molasses.

2.How do you prepare ethyl alcohol from ethylene?

Ethylene on addition of concentrated sulphuric acid form ethyl hydrogen sulphates, which on hydrolysis with water form ethyl alcohol.

$CH_2 = CH_2 + H_2SO_4 \longrightarrow CH_3 - CH_2 - HSO_4 \;\; CH_3 - CH_2 - OH$

Δ Ethyl alcohol

3.Give any three uses of ethyl alcohol.

(i) As a solvent for fats and many other organic compounds.

(ii) In the preparation of esters, used as perfumes.

(iii) In the manufacture of chemicals such as chloroform, chloral, iodoform, ether, acetic acid, ethylene, *etc*.

4.What is esterification? Give example.

When ethanol is warmed with ethanoic acid **in presence of a few drops** of concentrated sulphuric acid, sweet smell of ethyl ethanoate (ethyl acetate) is produced.

$$CH_3COOH + C_2H_5OH \longrightarrow \underset{\text{Ethylacetate}}{CH_3COOC_2H_5} + H_2O$$

The process of the formation of an ester (ethyl ethanoate) by the combination of ethanol with ethanoic acid is known as **esterification.**

5.Write down the classification of alcohols with one example each.

Alcohols are classified according to the number of hydroxyl groups present. The alcohols containing one, two, three or more hydroxyl groups are known as mono, di, tri or ployhydric alcohols respectively.

For example:

CH_3CH_2OH Ethyl alcohol (Monohydric)	CH_2OH | CH_2OH Ethylene glycol (Dihydric)	CH_2OH | $CHOH$ | CH_2OH Glycerol (Trihydric)

Monohydric Alcohols: Monohydric alcohols are classified as primary (1°), secondary (2°) and Tertiary (3°) alcohols depending upon whether the hydroxyl group is attached to a primary, secondary or tertiary carbon atom.

(a) When the carbon atom having the – OH group is attached to only one carbon atom, the alcohol is termed as a primary (1°) alcohol

$$\begin{array}{c} H \\ | \\ CH_3 - C^{1°} - OH \\ | \\ H \end{array} \qquad \text{Primary carbon}$$

Ethyl alcohol (primary alcohol, 1°)

(b) When the carbon atom having the – OH group is attached to two carbon atoms the alcohol is termed as secondary (2°) alcohol.

$$\begin{array}{c} CH_3 \\ | \\ CH_3 - C^{2°} - OH \\ | \\ H \end{array} \qquad \text{Secondary carbon}$$

Isopropylalcohol (secondary alcohol, 2°)

(c) When the carbon atom having the – OH group is attached to three carbon atoms, the alcohol is termed as tertiary (3°) alcohol.

$$\begin{array}{c} CH_3 \\ | \\ CH_3 - C^{3°} - OH \\ | \\ CH_3 \end{array} \qquad \text{Tertiary carbon}$$

t - Butyl alcohol(tertiary alcohol, 3°)

IV. Answer in detail.(T.B. Page 139)

1.How will you detect the presence of alcohol?

Detection of alcohol: Alcohol can be detected by the following tests.

1. **Sodium metal test:** when a small piece of sodium is added to an alcohol, hydrogen gas is evolved with effervescence.

$$2C_2H_5OH + 2Na \longrightarrow 2C_2H_5ONa + H_2\uparrow$$

Ethanol Sodium ethoxide

2. **Phosphorous pentachloride test:** On treatment with phosphorous pentachloride(PCl_5) alcohols become warm and hydrogen chloride gas is evolved.

$$C_2H_5OH + PCl_5 \longrightarrow C_2H_5Cl + POCl_3 + HCl\uparrow$$

2.What happens when (a) Ethanol is treated with acidic solution of potassium dichromate (b) Ethanol is burnt in air.

(a) Ethanol is oxidized to ethanoic acid by an acidic solution of potassium dichromate.

$$CH_3CH_2OH + 2[O] \xrightarrow[\text{heat}]{} \underset{\text{Ethanoic acid}}{CH_3COOH} + H_2O$$

(b) Ethanol burns in air with a blue flame to form carbon dioxide and water.

$$\underset{\text{Ethanol}}{C_2H_5OH} + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

3.Describe the Manufacture of ethanol from molasses.

Fermentation Process (From Molasses): Molasses is the mother liquor left over after the crystallization of sugar from the sugarcane juice. It is a dark coloured syrupy liquid containing 50 – 55 % total sugars (such as sucrose, glucose and fructose). **The manufacturing process involves the following steps:**

Step 1: Molasses is diluted with water to get 10% sugar solution.

Step 2: Dilute sulphuric acid is added to check the bacterial growth and to bring pH to 4 – 5 (Yeast thrives in acidic solution) and to stop the growth of unwanted bacteria.

Step 3: Some amount of ammonium sulphate or ammonium phosphate is added as food for yeasts.

Step 4: Yeast is added to the mixture and the solution is maintained at 30°C for 2 to 3 days.

The enzyme invertase present in yeast hydrolysis sucrose into glucose and fructose, the enzyme zymase present in yeast converts glucose and fructose to ethyl alcohol and carbon dioxide. **The process is called fermentation.**

$$\underset{\text{sucrose}}{C_{12}H_{22}O_{11}} + H_2O \longrightarrow \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$$

$$\underset{\text{Glucose or fructose}}{C_6H_{12}O_6} \longrightarrow \underset{\text{ethanol}}{2C_2H_5OH} + 2CO_2\uparrow$$

Yeast is filtered off. The resultant filtrate contains almost 10% of ethyl alcohol. It is also called **wash.**

Step 5: Wash is subjected to fractional distillation to get 95.5% alcohol and 4.5% of water. This is called rectified spirit. The rectified spirit cannot be further concentrated by fractional distillation as it form a constant boiling mixture or azeotropic mixture having its boiling point 78.13°C.

Step 6: Rectified spirit is mixed with quick lime (CaO) and allowed to stand for some time. It is then distilled. This is called lime of alcohol (98%). Final traces of water are removed by distillation with calcium metal to get absolute alcohol (100% alcohol). Anhydrous copper sulphate can also be used to remove final traces of water.

4. Give the IUPAC Name and common name for the following compounds.

(a) $CH_3 - \underset{\displaystyle OH}{\underset{|}{CH}} - \underset{\displaystyle CH_3}{\underset{|}{CH_2}} - CH_3$ **(c)** $CH_3 - CH - CH_2 - OH$

(b) $CH_3 - \overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{C}}}} - OH$ **(d)** $CH_3 - CH_2OH$

(a) **IUPAC name:** Butan-2-ol **Common name :** Sec-butyl alcohol

(b) **IUPAC name:** 2-Methyl propan-2-ol **Common name :** Tert-butyl alcohol

(c) **IUPAC name:** 2-Methyl propan-1-ol **Common name :** Isobutyl alcohol

(d) **IUPAC name:** Ethanol **Common name :** Ethyl alcohol

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1. Functional group of an alcohol is
   (a) – CO (b) – COOH (c) – OH (d) – CI
2. Ethyl alcohol is a
   (a) secondary alcohol (b) tertiary (c) primary alcohol (d) oxide group
3. Denatured spirit is ethanol mixed with
   (a) petrol (b) kerosene (c) water (d) pyridine
4. The formula for secondary butyl alcohol is

   (a) CH_3 CH CH_2 CH_3
   |
   OH

   (b) CH3 ╲ ╱ CH_3
   C
   H_3C ╱ ╲ OH

   (c) $CH_3CH_2CH_2OH$ (d) $CH_3CH_2CH_2OH$
5. Ethyl alcohol is
   (a) a beverage (b) explosives (c) antiseptic (d) poisonous gas.
6. Ethanol is the type of
   (a) secondary alcohol (b) tertiary alcohol (c) Primary alcohol (d) None of these
7. Ethyl alcohol is an example of an
   (a) aliphatic primary alcohol (b) aliphatic secondary alcohol
   (c) aliphatic tertiary alcohol (d) trihydric alcohol
8. Molasses is mainly a solution of
   (a) sucrose (b) ethyl alcohol (c) starch (d) maltose
9. The catalyst in the conversion of glucose and fructose to ethyl alcohol is
   (a) invertase (b) finely divided Nickel (c) zymase (d) V_2O_5
10. Which of the following reactions is used in the identification of ethyl alcohol?
    (a) esterification (b) reaction with Na (c) idoform reaction (d) all of 1,2,3
11. Butan – 2-ol is
    (a) primary alcohol (b) secondary alcohol (c) tertiary alcohol (d) none
12. Which of the following is not a characteristic of alcohols?
    (a) lower members ar insoluble in water and organic solvents but solubility regularly increases with molecular weight.
    (b) They are lighter than water
    (c) lower members have pleasant smell and burning taste, higher members are odourless and tasteless.
    (c) The boiling point increases with increase in molecular weight
13. Which does not contain ethanol?
    (a) absolute alcohol (b) carbinol (c) power alcohol (d) rectified spirit
14. Ethyl alcohol when treated with conc. H2SO4 gives
    (a) diethyl ether (b) ethyl hydrogen sulphate (c) ethylene (d) all of the above
15. Boiling point of alcohol is more than that of ether of corresponding molecular weight, because
    (a) alcohols are more soluble in water (b) ethers are non-polar molecules
    (c) hydrogen bonding exist between alcohol molecules (d) none of these
16. Absolute alcohol is prepared from rectified spirit by
    (a) steam distillation (b) azeotropic distillation
    (c) reduced pressure distillation (d) fractional distillation
17. Ethyl alcohol can be denatured by adding
    (a) acetone (b) ethanol (c) methanol (d) all of 1,2,3
18. Which of the following is a tertiary alcohol?
    (a) CH_3CH_2OH (b) $CH_3CHOHCH_3$

    (c)
    CH_3
    |
    CH_3 – C – CH_3
    |
    OH

    (d) CH_3 – $CHCH_2OH$
    |
    CH_3
19. The process that occurs when ethanol is converted into ethanal is

(a) reduction (b) hydrogention (c) oxidation (d) dehydration

20. Which is true of a compound with the formula $CH_3CH(OH)CH_3$
(a) It is a primary alcohol (b) it is a tertiary alcohol
(c) it can be oxidized to aldehyde (d) it can be oxidized a ketone

21. The – OH group of ethyl alcohol can be replaced by the action of
(a) H_2 (b) Cu (c) P/ I_2 (d) PCl_5

22. Power alcohol is
(a) an alcohol of 95% purity (b) a mixture of petrol hydrocarbons and ethanol
(c) rectified spirit (d) a mixture of methanol and ethanol

23. Rectified spirit is
(a) ethanol mixed with methanol (b) 50% ethanol + 50% water
(c) beonzoic acid (d) 95% ethanol +5% water

24. Denatured alcohol is
(a) rectified spirit (b) undistilled ethanol
(c) rectified spirit +5% methanol + naphtha + pyridine
(d) 50% ethanol + 50% methanol

25. Ethyl alcohol can be used for the preparation of
(a) ethylene (b) acetic acid (c) ethyl acetate (d) all these compounds

26. Which of the following is a primary alcohol?
(a) n- Pentyl alcohol (b) sec – Pentyl alcohol (c) tert – Pentyl alcohol (d) None of these

27. Ethanol can be obtained from all methods except
(a) Hydration of alkene (b) Hydration of Ethylene
(c) Reduction of aldehyde / ketones with Zn – Hg / HCl (d) Fermentation of Molasses

28. The reagent used for oxidation of an Ethanol is
(a) $K_2Cr_2O_7$ (b) Calcium chloride (c) Al_2O_3 (d) NaCl

29. Which one of the following on oxidation gives carboxylic acids?
(a) Primary alcohol (b) Secondary alcohol (c) Tertiary alcohol (d) All of these

30. What is formed when a primary alcohol undergoes catalytic oxidation?
(a) Aldehyde (b) Ketone (c) Alkene (d) Amine

31. During the fermentation of molasses the enzyme that converts glucose to alcohol is
(a) Maltase (b) Invertase (c) Zymase (d) Oxidase

32. Denatured alcohol is a mixture of ethyl alcohol and
(a) Methanol and Toluene (b) Methanol and pyridine
(c) Methanol and acetic acid (d) None of the above is correct

33. Rectified spirit is a mixture of
(a) 95% C_2H_5OH and 5% H_2O (b) 94% C_2H_5OH and 6% H_2O
(c) 95.87% C_2H_5OH and 4.13% H_2O (d) 94.47% C_2H_5OH and 5.53% H_2O

34. Power alcohols is
(a) An alcohol of 95% purity (b) Absolute alcohol
(c) Rectified spirit (d) A mixture of petrol and ethanol

35.Fermentation is a / an
(a) Fast process (b) slow process (c) Reversible process (d) None of these

36.Which of the following enzymes converts starch into maltose?
(a) Zymase (b) Maltase (c) Diastase (d) invertase

37.Absolute alcohol can be obtained from rectified spirit by
(a) Fractional distillation (b) distillation (c) Vacuum distillation (d) Steam distillation

38.The reaction between an alcohol and an acid with the elimination of water molecule is called?
(a) Elimination (b) Esterification (c) Saponification (d) Etherification

39.Ethyl alcohol is denatured by?
(a) Methanol and formic acid (b) KCN
(c) CH_3OH and C_6H_6 (d) CH_3OH and pyridine

40. Methyl alcohol on oxidation with acidified $K_2Cr_2O_7$ gives?
(a) CH_3COCH_3 (b) CH_3CHO (c) $HCOOH$ (d) CH_3COOH

41. When ethyl alcohol and acetic acid is heated with conc. H_2SO_4, the product obtained is?
(a) $CH_3COOC_2H_5$ (b) C_2H_6 (c) C_2H_4 (d) C_2H_2

42. Ethyl alcohol is manufactured on an industrial scale by the fermentation of?
(a) $C_6H_{12}O_6$ (b) $C_{11}H_{22}O_{11}$ (c) Molasses (d) CH_3COOH

43. Fermentation of sugar with yeast gives?
(a) CH_3OH (b) $HCHO$ (c) C_2H_5OH (d) $CH_3CH_2CH_2OH$

44. The enzyme which converts glucose and fructose into ethyl alcohol is
(a) Diastase (b) Invertase (c) Zymase (d) Maltase

45. Alcohol is commonly used as a:
(a) Preservative (b) Antifreezing compound (c) Beverage (d) All
46.Ethyl alcohol can be used for the preparation of?
(a) Ester (b) Ethylene (c) Acetic acid (d) All
47.C_2H_5OH on oxidation with $K_2Cr_2O_7$ gives?
(a) Acetaldehyde (b) Acetic acid (c) Formaldehyde (d) Formic acid
48. Denatured spirit is mainly used as?
(a) Fuel (b) Solvent in preparing varnishes (c) Material in preparing beverages (d) Wine
49. Which gas is evolved during fermentation?
(a) CO (b) CO_2 (c) H_2 (d) CH_4
50. On hydration, ethylene gives?
(a) Ethanol (b) Ethylene glycol (c) Ethyl alcohol (d) methyl alcohol
51. Ethanol gives chloroethane on treatment with
(a) PCl_5 (b) PCl_3 (c) $SOCl_2$ (d) All
52. 95% ethanol is called?
(a) Power alcohol (b) Absolute alcohol (c) Rectified spirit (d) Methylated spirit
53.denaturated spirit contains about?
(a) 48% methanol by weight (b) 10% methanol by weight
(c) 5% methanol by weight (d) 90% methanol by weight
54. $A \xrightarrow[H_2SO_4]{K_2Cr_2O_7} CH_3COOH$ What is A?
(a) C_2H_5OH (b) Isopropyl alcohol (c) Propyl alcohol (d) methyl alcohol
55. Which is used as antifreeze compound?
(a) Methanol (b) Ethane (c) Methane (d) All
56. Denatured alcohol is?
(a) Rectified spirit (b) Undistilled ethanol
(c) Rectified spirit + methanol + pyridine (d) 50% ethanol + 50% methanol
57. Fermentation is an
(a) Endothermic (b) Exothermic (c) Reversible reaction (d) None
58. Sodium acetate on decarboxylation gives ....
(a) methane (b) ethane (c) propane (d) butane
59. The general formula of primary alcohol is:
(a) > CHOH (b) $-\overset{|}{\underset{|}{C}}-OH$ (c) — CH_2OH (d) > C $(OH)_2$
60. Which one is primary alcohol?
(a) Buten – 2 –ol (b) Propan –2 –ol (c) Butan –1-ol (d) 2, 3, -Dimethylhexane –4 –ol
61. Ethyl alcohol is industrially prepared from ethylene by
(a) Permanganate oxidation (b) catalytic reduction
(c) Absorbing in H_2SO_4 followed by hydrolysis (d) Fermentation
62. Ethanol containing some methanol is called
(a) Absolute spirit (b) Rectified spirit (c) Power alcohol (d) Methylated spirit
63. The enzyme which can catalyse the conversion of glucose to ethanol is
(a) Zymase (b) Invertase (c) Maltase (d) Diastase
64. Alcohol fermentation is brought about by the action of
(a) CO_2 (b) O_2 (c) Invertase (d) Yeast
65.Which is used as an antifreeze?
(a) Glycol (b) ethyl alcohol (c) Water (d) Methanol

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

II. Answer in one or two sentences:

1. What are alcohols? Give the general formula.

Alcohols are organic compounds with general formula R – OH where R is an alkyl group. These may be considered as hydroxy derivatives of hydrocarbons in which one or more hydrogen atoms are replaced by hydroxyl (– OH) groups. In this chapter, we will be dealing with only aliphatic alcohols.

2. Define and give example for 1°, 2°, 3° alcohols.

(i) When the carbon atom having the – OH group is attached to only one carbon atom, the alcohol is termed as a primary (1°) alcohol

$$\begin{array}{c} \mathrm{H} \\ | \\ \mathrm{CH_3 - C^{1^\circ} - OH} \\ | \\ \mathrm{H} \end{array} \qquad \text{Primary carbon}$$

Ethyl alcohol (primary alcohol, 1°)

(ii) When the carbon atom having the – OH group is attached to two carbon atoms the alcohol is termed as secondary (2°) alcohol.

$$\begin{array}{c} \mathrm{CH_3} \\ | \\ \mathrm{CH_3 - C^{2^\circ} - OH} \\ | \\ \mathrm{H} \end{array} \qquad \text{Secondary carbon}$$

Isopropylalcohol (secondary alcohol, 2°)

(iii) When the carbon atom having the – OH group is attached to three carbon atoms, the alcohol is termed as tertiary (3°) alcohol.

$$\begin{array}{c} \mathrm{CH_3} \\ | \\ \mathrm{CH_3 - C^{3^\circ} - OH} \\ | \\ \mathrm{CH_3} \end{array} \qquad \text{Tertiary carbon}$$

t - Butyl alcohol(tertiary alcohol, 3°)

3. The IUPAC name of alcohol is derived from alkane- Explain by giving examples.

(a) **The IUPAC name of the alcohol is obtained by replacing ‘e’ of the corresponding alkane by the suffix ‘–ol’.**

Example: $\underset{\text{Methane}}{CH_4} \longrightarrow \underset{\text{Methanol}}{CH_3OH}$ $\quad \underset{\text{Ethane}}{C_2H_6} \longrightarrow \underset{\text{Ethanol}}{C_2H_5OH}$

(b) The position of the - OH group is indicated by numbering the carbon chain so as to give the lowest possible number to the carbon bearing the – OH group,

Example:

(1) $\overset{3}{CH_3} - \overset{2}{CH_2} - \overset{1}{CH_2} - OH \longrightarrow$ Propan -1-ol

(2) $\overset{3}{CH_3} - \underset{\underset{OH}{|}}{\overset{2}{CH}} - \overset{1}{CH_3} \longrightarrow$ Propan -2-ol

4. List the physical properties of ethanol.

1. Ethanol is a colourless liquid having pleasant smell.
2. Ethanol boils at 351 K.
3. It is miscible with water in all proportions.
4. It is nonconductor of electricity because it does not contain ions.

5. What happens when ethanol is heated with acidified potassium dichromate?

Ethanol is oxidized to ethanoic acid by an acidic solution of potassium dichromate.

$$CH_3CH_2OH + 2[O] \xrightarrow[\text{heat}]{} \underset{\text{Ethanoic acid}}{CH_3COOH} + H_2O$$

6. What are monohydric alcohols? Give two examples.

Alcohols containing one-OH group in a molecule are called monohydric alcohols. e.g., methyl alcohol, ethyl alcohol.

7. What are dihydric alcohol? Give example.

Alcohols containing two-OH groups in a molecule are called dihydric alcohols. e.g., glycol, $CH_2OH - CH_2OH$.

8. What are trihydric alcohol? Give one example.

Alcohols containing three-OH groups in a molecule are called trihydric alcohols.

e.g., CH_2OH
|
$CHOH$ Glycerol
|
CH_2OH

9. How do you classify monohydric alcohols?

The monohydric alcohols are classified as (a) Primary alcohols (1^o), (b) Secondary alcohols (2^o) (c) Tertiary alcohols (3^o).

10. What is a primary alcohol? Give example.

A primary alcohol is one in which the carbon atom attached to the-OH group is in turn connected to one other or no other carbon atom.

e.g., Ethyl alcohol, CH_3-CH_2OH.

11. What is a secondary alcohol? Give example.

A secondary alcohol is one in which the carbon atom attached to the –OH group is in turn connected to two other carbon atoms.

e.g., Isopropyl alcohol, $CH_3 - CHOH - CH_3$.

12. What is a tertiary alcohol? Give example.

A tertiary alcohol is one in which the carbon atom attached to the –OH group is in turn connected to three other carbon atoms.

OH
|
e.g., Tertiary butyl alcohol, $CH_3 - C - CH_3$
|
CH_3

13. What is the trivial system of naming of alcohols? Give an example.

In the trivial system, saturated aliphatic alcohols are named as alkyl alcohols.

e.g., $\underset{\text{Methane}}{CH_4} \longrightarrow \underset{\text{Methyl}}{CH_3\text{-}} \longrightarrow \underset{\text{Methyl alcohol.}}{CH_3\text{-}OH}$

14. What is lime of alcohol?

Rectified spirit is mixed with quick lime (CaO) and allowed to stand for some time. It is then distilled. **This is called lime of alcohol (98%).**

15. What is wash?

In fermentation process, the resultant filtrate containing 10% ethl alcohol is called wash.

16. What is rectified spirit?

Wash is subjected to fractional distillation to get 95.5% alcohol and 4.5% of water. This is called rectified spirit.

17. What is the IUPAC system of naming alcohols? Give an example.

(a) The IUPAC name of the alcohol is obtained by replacing 'e' of the corresponding alkane by the suffix '-ol'.

$\underset{\text{Methane}}{CH_4} \longrightarrow \underset{\text{Methanol}}{CH_3OH}$

(b) The position of the –OH group is indicated by numbering the carbon chain so as to give the lowest possible number to the carbon bearing –OH group.

3 2 1

e.g., $CH_3 - CH_2 - CH_2OH$
Propan-1-o1

III. Answer in brief:

1.Explain the trival system of naming of alcohols by giving three examples.

In the trivial (common system) system, saturated aliphatic alcohols are named as alkyl alcohols.

Example:

CH_4 $\longrightarrow$	CH_3^- $\longrightarrow$	$CH_3 - OH$
Methane	Methyl	Methyl alcohol
C_2H_6 $\longrightarrow$	$C_2H_5 -$ $\longrightarrow$	$C_2H_5 - OH$
Ethane	Ethyl	Ethyl alcohol

The prefixes like normal (n), iso and tertiary are used in the case of isomeric alcohols.

Example:

(1) $CH_3 - CH_2 - CH_2 - OH$ *n* - propyl alcohol

(2) $CH_3 - CH - OH$ isopropyl alcohol
|
CH_3

(3) $CH_3 - CH - CH_2 - OH$ isobutyl alcohol $CH_3 - CH -$ isopropyl group
|
CH_3

9 – ETHERS

INTRODUCTION

The word ether is derived from Greek word **aither** meaning the clear sky or air; it represents the airy or volatile nature of typical compounds of this class. Ethers are the compounds having the general formula R – O – R. They can be regarded as the dialkyl derivatives of water or monoalkyl derivatives of alcohols.

$$\underset{}{H-O-H} \xrightarrow[+2R]{-2H} \underset{\text{Ether}}{R-O-R} \quad ; \quad \underset{\text{Alcohol}}{R-O-H} \xrightarrow[+R]{-H} \underset{\text{Ether}}{R-O-R}$$

The two R groups attached to oxygen may be identical or different. Ethers in which the two R groups are identical are known as **symmetrical** or **simple** ethers while those in which the two groups are different are called **unsymmetrical** or mixed ethers. Thus we have:

CH_3-O-CH_3	$CH_3-O-C_2H_5$
Dimethyl ether	Ethyl methyl ether
(Symmetrical)	(Unsymmetrical)

Ethers are further classified into the following two categories:

1. Aliphatic ethers. In such ethers, the R and R' are both alkyl groups. For example:

CH_3-O-CH_3	$CH_3CH_2-O-CH_3$
Dimethyl ether	Ethyl methyl ether

2. Aromatic ethers. In such ethers one or both R and R' are aryl groups. For example:

$C_6H_5-O-C_6H_5$	$CH_3-O-C_6H_5$
Diphenyl ether	Methyl phenyl ether

IMPORTANT TERMS & DEFINTIONS

Ethers: They are organic compounds with the general formula ($R-O-R^1$) where R and R^1 are alkyl or aryl groups. The groups R and R^1 may be either the same or different.

Types of Ethers: There are two types of ethers:(i) Simple ethers or symmetrical ethers, (ii) Mixed ethers or unsymmetrical ethers. If Rand R^1 are the same the ethers are called **'simple ethers'**. and if R and R^1 are different these are called **'mixed ethers'**.

Nomenclature: (a) Common name (b) IUPAC name

(a) Common system: In the Common system the ethers are named according to the alkyl group bonded to the oxygen atoms. The two-alkyl groups bonded to the functional group (-O-) are written alphabetically followed by the word ether **For example CH_3 - O - C_2H_5**
i.e. Ethyl methyl ether (arranged alphabetically)

If both the groups are similar prefix like '–di' is attached.
For example CH_3 - O - CH_3 i.e Dimethyl ether

(b) IUPAC System: In IUPAC System, the ethers are named as alkoxy alkanes. The oxygen atom is takes with the smaller alkyl group while the larger alkyl group forms the parent chain

For example CH_3 - O - C_2H_5	CH_3 - O - $\overset{1}{C}H_2-\overset{2}{C}H_2-\overset{3}{C}H_3$
i.e.1- Methoxy ethane	i.e. Methoxy propane

Williamson's ether synthesis: When an alkyl halide is heated with sodium or potassium alkoxide, an ether is obtained. Both symmetrical and unsymmetrical ethers can be prepared by this method.

1. Physical state: Dimethyl ether and ethyl methyl ether are gases at room temperatures. Other lower members are colourless liquids, which are highly volatile.

2. Boiling points: Ethers have much lower boiling points compared to isomeric alcohols as they are not associated with hydrogen bonds. Their boiling points are comparable to the corresponding alkanes.

3.Volatility and flammability: Due to low boiling points, the lower members are highly volatile and thus catch fire immediately. So lower ethers are highly inflammable.

4.Solubility: Ethers are soluble in hydrocarbons and other non-polar solvents. Ethers are generally insoluble in water, but their solubility in water is not negligible.

5. Inertness: Owing to the absence of active groups and multiple bonds, ethers are comparatively inert substances. They are not easily attacked by alkalies, dilute acids, PCl_5. metallic sodium etc. They undergo chemical reactions under specific conditions. Some of the reactions of ethers are due to:

SELF EVALUATION (T.B. Page 144)

I. Choose the correct answer.

1. The .... nature, has made ether a versatile solvent.

   (a) Inert (b) High reactive

   (c) Easily boiling (d) None

2. Ether on reaction with chloride in dark gives ....

   (a) α, α'-Dichlorodiethyl ether (b) Perchloro diethyl ether

   (c) Ethanol (d) Ethyl chloride

3. Ether when heated with PCl_5 gives ....

   (a) C_2H_5OH (b) C_2H_5Cl

   (c) C_2H_5I (d) C_2H_6

4. Diethyl ether when treated with excess HI .... are formed.

   (a) Ethyl iodide only (b) ethanol + ethyl iodide

   (c) ethyl iodide + H_2O (d) None

5. Ether is used as an ....

   (a) Anaesthetic (b) Antiseptic

   (c) Anti pyretic (d) all the above

Answers:

1. (a) 2. (a) 3. (b) 4. (c) 5. (a)

II. Answer the following in One or Two sentences.(T.B. Page 145)

1.Give the common name for the following:

(a) $CH_3 - O - CH_3$; (b) $CH_3 - O - C_2H_5$

(a) Dimethyl ether (b) Ethyl methyl ether

2.Complete: $C_2H_5 - O - C_2H_5 + Cl_2 \longrightarrow$ Complete the reaction?

$$CH_3 - CH_2 - O - CH_2 - CH_3 + Cl_2 \longrightarrow CH_3 - \underset{\displaystyle Cl}{\underset{|}{CH}} - O - \underset{\displaystyle Cl}{\underset{|}{CH}} - CH_3 + 2HCl$$

α, α' - Dichlorodiethyl ether

3.Ether is a versatile solvent and medium for reaction. – Give reason.

Ethers are able to dissolve large variety of organic compounds. The inert character has made ether a versatile solvent and medium for reaction.

4.$C_2H_5 - O - C_2H_5 + PCl_5 \longrightarrow$? Complete and balance the equation

$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5} + PCl_5 \longrightarrow \underset{\text{Ethyl chloride}}{2C_2H_5Cl} + POCl_3$$

5.$C_2H_5 - O - C_2H_5 + HI \longrightarrow$? Complete and balance the equation

$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5} + HI \longrightarrow \underset{\text{Ethyl alcohol}}{C_2H_5OH} + \underset{\text{Ethyl iodide}}{C_2H_5I}$$

III. Answer in brief. (T.B. Page 145)

1.What are the uses of ether?

1. Used as a refrigerant.
2. Used as a solvent for oils, gums and resins.
3. Used as an anaesthetic in surgery.
4. Used as a freezing mixture in the form of ether and dry ice.

2.Give the IUPAC for the following.

(a) $CH_3OCH_2CH_2CH_3$ (b) $CH_3CH_2CH_2 - O - CH_2CH_3$ (c) $CH_3 - O - CH_3$

IUPAC names:

(a) Methoxy propane (b) Ethyoxy propane (c) Methoxy methane

3.What is Williamson's ether synthesis?

When an alkyl halide is heated with sodium or potassium alkoxide, an ether is obtained. Both symmetrical and unsymmetrical ethers can be prepared by this method.

$$\underset{\text{Sodium ethoxide}}{C_2H_5ONa} + \underset{\text{Methyl bromide}}{BrCH_3} \longrightarrow \underset{\text{Ethyl methyl ether}}{C_2H_5OCH_3} + NaBr$$

$$\underset{\text{Sodium Ethoxide}}{C_2H_5ONa} + \underset{\text{Ethyl Bromide}}{BrC_2H_5} \longrightarrow \underset{\text{Diethyl Ether}}{C_2H_5-O-C_2H_5} + Na\,Br$$

4.Diethyl ether + excess HI $\rightarrow$? $\rightarrow$? Complete and balance the equation?

When ether is heated with **excess of concentrated hydroiodic acid** alkyl halides are formed.

$$\underset{\text{Diethyl ether}}{C_2H_5-O-C_2H_5} + 2HI \longrightarrow \underset{\text{Ethyl iodide}}{2C_2H_5I} + H_2O$$

5.Mention the types of ethers. Give example?

There are two types of ethers.

(i) Simple ethers or symmetrical ethers, *(ii)* Mixed ethers or unsymmetrical ethers.

If Rand R′ are **the same** the ethers are called **'simple ethers'**. and if R and R′ are different these are called **'mixed ethers'**.

Simple ether $C_2H_5–O–C_2H_5$ Mixed ether $CH_3–O–C_2H_5$

Diethyl ether **Ethyl methyl ether**

IV. Answer in detail. (T.B. Page 146)

1.(a) Give the common name of the following:

1. $CH_3-O-CH_2CH_2CH_3$ 2. $CH_3CH_2CH_2-O-CH_2CH_3$

Common names: 1. Methyl propyl ether, 2. Ethyl propyl ether.

(b) Write a note on reactions due to the alkyl group in ethers.

Halogenation: Ethers react with chlorine or bromine in the dark to give substituted products at a-carbon atoms.

$$CH_3-CH_2-O-CH_2-CH_3 + Cl_2 \longrightarrow CH_3-\underset{\underset{Cl}{|}}{CH}-O-\underset{\underset{Cl}{|}}{CH}-CH_3 + 2HCl$$

α, α′ - Dichlorodiethyl ether

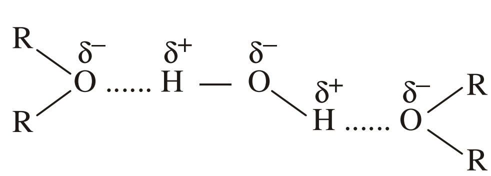
In the presence of sunlight, all the hydrogens of the ether are substituted by halogen atoms.

$$\underset{\text{Diethyl ether}}{C_2H_5-O-C_2H_5} + 10Cl_2 \longrightarrow \underset{\text{Perchloro diethyl ether}}{C_2Cl_5-O-C_2Cl_5} + 10HCl$$

2.Write a note on solubility of ethers.

Ethers are soluble in hydrocarbons and other non-polar solvents. Ethers are generally insoluble in water, but their solubility in water is not negligible. *For example,* the solubility of diethyl ether in water is 8g in 100g of water at 293K.

The solubility of lower ethers in water is **due to the formation of hydrogen bond** between water and ether compounds as shown below.



This is due to some hydrogen bonding between water and ether molecules.

Ethers are able to dissolve large variety of organic compounds. The inert character has made ether a versatile solvent and medium for reaction. When inhaled, ether vapour produces unconsciousness and insensibility to pain. It is therefore, used as a general anaesthetic.

3.Explain the chemical properties of ether involving the cleavage of C – O bond.

(i) Reaction with dilute sulphuric acid: When ether is heated with dilute sulphuric acid under pressure alcohols are formed.

$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5 + H_2O} \xrightarrow[\text{Pressure}]{\Delta} \underset{\text{Ethyl alcohol}}{2C_2H_5OH}$$

(ii) Reaction with phosphorus Pentachloride:

When an ether is heated with phosphorus pentachloride, alkyl halides are formed.

$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5} + PCl_5 \longrightarrow \underset{\text{Ethyl chloride}}{2C_2H_5Cl} + POCl_3$$

(iii) Reaction with hydroiodic acid:

When ethers heated with **conc. Hydroiodic acid**, an alcohol and an alkyl halide are formed.

$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5} + HI \longrightarrow \underset{\text{Ethyl alcohol}}{C_2H_5OH} + \underset{\text{Ethyl iodide}}{C_2H_5I}$$

When ether is heated with **excess of concentrated hydroiodic acid** alkyl halides are formed.

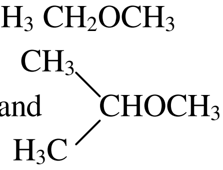
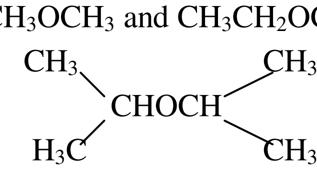
$$\underset{\text{Diethyl ether}}{C_2H_5 - O - C_2H_5} + 2HI \longrightarrow \underset{\text{Ethyl iodide}}{2C_2H_5I} + H_2O$$

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1. Which of the following is a simple ether?
   (a) $C_2H_5OCH_3$ (b) CH_3OCH_3 (c) $C_6H_5OCH_3$ (d) All are simple ether.
2. Which of the following compounds is a mixed ethers?
   (a) $CH_3COC_2H_5$ (b) $C_2H_5OC_2H_5$ (c) CH_3OCH_3 (d) $C_2H_5OC_2H_5$
3. IUPAC name of ethyl di ether is
   (a) Ethoxyethane (b) 3 – Ethoxypropane (c) Methyoxy methane (d) Ethoxypropane
4. IUPAC name of ethyl methyl ether is
   (a) Propane (b) methoxypropane (c) Methoxy ethane (d) Ethoxy methane
5. The solubility of lower ether's in water due to formation of
   (a) H-bonds (b) Covalent bonds (c) Dipole – dipole forces (d) Ion- dipole forces.
6. The boiling point of acetic acid is ……………. ^oC
   (a) 118 (b) 181 (c) 811 (d) 188
7. Methyl ethers cab be prepared by reaction of alcohols with
   (a) Alkyl halides (b) Diazomethane (c) Grignard reagent (d) None of these.
8. Reaction between sodium exthoxide and bromoethane yields
   (a) Methyl ethyl ether (b) Dimethyl ether (c) Diethyl ether. (d) Propane.
9. Which one is formed when sodium ethoxide is heated with ethyl iodide?
   (a) Phenetole (b) Ethyl Methyl ether (c) Diethyl ether (d) none of these
10. Which of the following reactions does not yield an alkyl halide?
   (a) Diethyl ether (b) Diethyl ether + HI (c) Diethyl ether + PCl_5 (d) Diethyl ether + HCl
11 An example of a compound with a functional group – O – is?
   (a) Acetic acid (b) Ethanol (c) diethyl ether (d) Methyl acetate
12. Diethyl ether on treatment with Cl_2 in presence of sunlight give?
   (a) Trichlorodiethyl ether (b) Perchlorodiethyl ether (c) anisole (d) ethane
13. Which of the following is simple ether?
   (a) $C_2H_5OC_2H_5$ (b) CH_3OCH_3 (c) $C_6H_5OC_6H_5$ (d) All are simple ethers
14. Which of the following compounds is mixed ethers?
   (a) $C_3H_7OC_3H_7$ (b) $C_2H_5OC_2H_5$ (c) CH_3OCH_3 (d) $CH_3OC_2H_5$
15. The common name for CH_3OCH_3
   (a) Dimethyl ether (b) Methoxy methane (c) Methyl ether (d) methoxy ethane
16. IUPAC name of Diethyl ether is
   (a) Ethoxy ethane (b) 2-Ethoxypropene (c) 1-Ethoxy methane (d) methoxy ethane
17. IUPAC name of CH_3-CH(Cl)-O-CH(Cl)-CH_3 is
   (a) α, α^1 Dichloro diethyl ether (b) α, α^1 Dichloro dimethyl ether
   (c) α, α^1 chloro diethyl ether (d) α, α^1 Dichloro methyl ethyl ether
18. Ether which is liquid at room temperature is
   (a) $C_2H_5OCH_3$ (b) CH_3OCH_3 (c) $C_2H_5OC_2H_5$ (d) None
19. The solubility of lower ethers in water is due to formation of

(a) H-bonds (b) Covalent bonds (c) Ionicbond (d) Ion-dipole forces

20. Williamson synthesis involves to prepare
(a) Symmetrical ethers only (b) Both symmetrical and unsymmetrical ethers
(c) unsymmetrical ethers (d) None of the abow
21. Ether is obtained from sodium ethoxide
(a) in presence Methyl bromide (b) in presence of H_2SO_4 at 474 K
(c) in presence of H_2SO_4 (d) in presence of Ethanol
22. Methyl ethers can be prepared by reaction of sodium ethoxide is
(a) Alkyl halides (b) Diazomethane (c) Grignard reagent (d) None of these
23. Reaction between sodium ethoxide and bromoethane yields
(a) Methyl ethyl ether (b) Dimethyl ether (c) Diethyl ether (d) Propane
24. Number of chlorine atoms in per chloro diethyl ether is …………..
(a) two (b) five (c) ten (d) twelve
25. Ethers are …………..
(a) acidic (b) weakly basic (c) neutral (d) amphoteric
26. ………………. Vapours produce unconsciousness.
(a) Alcohol (b) Ether (c) Ester (d) aldehydes
27. ………………… form explosive mixture with air.
(a) Alcohols (b) Esters (c) ether (d) aldehydes
28. Ethyl alcohol is an isomer of …………….
(a) diethyl ether (b) dimethyl ether (c) Ethers (d) Aldehydes
29. The formula of sodium ethoxide is …………..
(a) C_2H_5Na (b) C_2H_5ONa (c) CH3ONa (d) NaOH
30. Ethers are isomeric with …………………
(a) esters (b) aldehydes (c) alcohols (d) ketones
31. $C_2H_5 - O - C_2H_5 + A \longrightarrow C_2H_5OH + C_2H_5I$. A is ……………….
(a) HI (b) I_2 (c) CH_3I (d) H_2 and I_2
32. $C_2H_5 - O - C_2H_5 + A \longrightarrow C_2H_5Cl + POCl_3$. A is ……………….
(a) CH_3Cl (b) PCl_5 (c) P and Cl_2 (d) Cl_2
33. $C_2H_5 - O - C_2H_5 + H_2O \xrightarrow[H_2SO_4]{dil}$ The product is ……………..
(a) $C_2H_5OCH_3$ (b) H_2O_2 (c)2 $C_2H_5\,OH$ (d) $C_2H_5\ H_2\ SO_4$
34. Identify the pairs of ethers, known as mixed ethers …………………
(a) CH_3OCH_3 and $CH_3\ CH_2OCH_3$ (b) CH_3OCH_3 and $CH_3CH_2OCH_2CH_3$
(c) $CH_3\ CH_2\ OCH_3$ and  (d) 
35. The one which is not a simple ether is ……………
(a) $CH_3\ O\ CH_3$ (b) $CH_3CH_2OCH_2CH_3$ (c) $CH_3OCH_2CH_3$ (d) both (a) and (b)

Answers:

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

II. Answer in one or two sentences:

1. Give the general formula of ethers.

The general formula of ethers is R-O-R’ where R and R’ are alkyl groups.

2. Give two examples of simple ethers.

Examples of two simple ethers are (a) diemethyl ether, (b) diethyl ether.

3. Give two examples of mixed ethers.

Examples of two mixed ethers are (a) ethyl methyl ether, (b) methyl propyl ether.

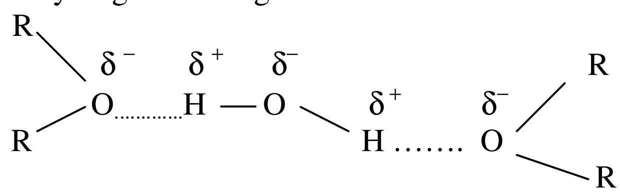
4. Write the IUPAC name of CH_3-O-CH-CH_3

$$\begin{array}{l} CH_3\text{-O-}\underset{\displaystyle CH_3}{\underset{|}{CH}}\text{-}CH_3 \end{array}$$

The IUPAC name of $CH_3\text{-O-}\overset{2}{\underset{\underset{\displaystyle {}_{1}CH_3}{|}}{CH}}\text{-}\overset{3}{CH_3}$ is 2-methoxy propane.

5. Draw a neat structure to show the hydrogen bonding between water and ether.

The hydrogen bonding in ether with water is as shown below.



6. Give the action of air on ethers?

Vapours of ether form explosive mixture with air due to formation of **peroxide**.

$$2R - O\text{-} R' + O_2 \longrightarrow 2R\text{-O-O-}R'.$$

7. Write the equation for the action of chlorine on diethyl ether in presence of sunlight.

$$C_2H_5\text{-O-}C_2H_5 + 10Cl_2 \xrightarrow{\text{Sunlight}} \underset{\text{Perchloro diethyl ether}}{C_2Cl_5\text{-O-}C_2Cl_5} + 10HCl$$

8. Name the various types of reactions of ethers.

The important types of reactions of ethers are

(a) Reactions of the alkyl group.

(b) Reactions due to cleavage of the C –O bond.

(c) Reactions due to lone pair of electrons on oxygen.

9.Give the common and IUPAC name for:

(a) $CH_3 - O - C_2H_5$; (b) $CH_3 - O - CH_3$;

Compound	Common name	IUPAC name
(a) $CH_3 - O - C_2H_5$	Ethyl methy ether	Methoxy ethane
(b) $CH_3 - O - CH_3$	Dimethyl ether	Methoxy methane

10. Write a short note on IUPAC system of naming ethers.

In IUPAC System, the ethers are named as alkoxy alkanes. The oxygen atom is taken with the smaller alkyl group while the larger alkyl group forms the parent chain.

For example, $\underset{\text{(Small group)}}{\underset{\uparrow}{CH_3}} - O - \underset{\text{(Larger group)}}{\underset{\uparrow}{C_2H_5}}$ *i.e.* Methoxy ethane

$\underset{\text{(Small group)}}{\underset{\uparrow}{CH_3}} - O - \overset{1}{CH_2} - \underset{\text{(Larger group)}}{\underset{\uparrow}{\overset{2}{CH_2}}} - \overset{3}{CH_3}$ *i.e.* Methoxy propane

11. Ethers have low boiling point than alcohols – why?

Ethers have much lower boiling points compared to isomeric alcohols as they are not associated with hydrogen bonds. Their boiling points are comparable to the corresponding alkanes.

12. Lower member of ethers are highly inflammable- why?

Due to low boiling points, the lower members are highly volatile and thus catch fire immediately. So lower ethers are highly inflammable. Like methane their vapours form explosive mixture with air.

13. Why ethers are inert substances?

Inertness: Owing to the absence of active groups and multiple bonds, ethers are comparatively inert Substances. So, they are not easily attacked by alkalies, dilute acids, PCl_5, metallic sodium *etc*. They undergo chemical reactions under specific conditions.

Answer in brief:

1. How do you convert diethyl ether into ethyl alcohol?

When diethyl ether is heated with dilute sulphuric acid under pressure, ethyl alcohol is formed.

$$\underset{\text{Diethy ether}}{C_2H_5\text{-}O\text{-}C_2H_5} + H_2O \xrightarrow[\Delta \\ \text{Pressure}]{\text{Dil. } H_2SO_4} \underset{\text{ethyl alcohol}}{2C_2H_5OH}$$

ANYTHING TO BE ADDED

10 – CARBONYL COMPOUNDS

INTRODUCTION

ALIPHATIC ALDEHYDES AND KETONES

Aldehydes and ketones are compounds containing carbonyl ($>C = O$) group. In aldehydes, the carbonyl group is linked either to two hydrogen atoms or to one hydrogen atom and one alkyl group. In ketones, the carbonyl group is linked to two alkyl groups.

Aldehydes and ketones may be represented as

$$R - \overset{\overset{H}{|}}{C} = O$$ (where R = H or any alkyl group)

Aldehyde

$$R - \overset{\overset{O}{\|}}{C} - R^1$$ (where R and R^1 may be same or different alkyl groups)

Ketone

If the two-alkyl groups ($R \& R^1$) are different, the ketone is said to be a **mixed ketone** and if R and R^1 represent the same alkyl group, the ketone is referred to as a **simple ketone**.

The functional group of aldehyde is $\overset{H}{\diagdown}\!\!\!>C = O$ and is called the aldehydic group while the functional group of ketones is $>C = O$ and is called the ketonic group.

Since both aldehydes and ketones contain the same carbonyl group, they have many similar chemical properties. But aldehydes differ from ketones in many respects due to the presence of a hydrogen atom on the carbonyl group of aldehydes.

IMPORTANT TERMS & DEFINITIONS

Carbonyl Compounds: Organic compounds containing carbonyl ($>C = O$) group are known as carbonyl compounds.

Nomenclature of Aldehydes

Common system: Aldehydes are named after the carboxylic acids they form on oxidation. The name is obtained by replacing the terminal 'ic acid' in the name of the acid by the suffix *aldehyde.*

IUPAC system: The ending of the name (suffix) of aldehyde is **'al'**. The names of aliphatic aldehydes are derived from the name of the corresponding alkane by replacing the terminal **'e'** by the suffix **'al'**.

Name of aldehyde = Name of corresponding alkane -e + al.

Nomenclature of Ketones

***Common system*:** The names of ketones are obtained by naming the two alkyl groups attached to the keto group (alphabetically) and adding the suffix 'ketone'

***IUPAC system*:** The characteristic ending for ketones is -'one'. The names of individual aliphatic ketones are derived by replacing the terminal 'e' in the name of the corresponding alkane by the suffix -'one'.

Carboxylic acids: Organic compounds which contain the carboxyl functional group (-COOH) are called the **carboxylic acids. Their general formula is** $R - \underset{\underset{O}{\|}}{C} - OH$ **or R -COOH where R is an alkyl group.**

Fatty acids: The long chain monocarboxylic acids are commonly called Fatty acids because many of them are obtained by the hydrolysis of animal fats or vegetable oils. Eg. Stearic acid, Palmitic acid.

Nomenclature of Monocarboxylic Acids

Common system: The common names are usually derived from the Latin or Greek word that indicates the original source of the acid. For example: Formic acid is present in ants (Latin formica = ants) and acetic acid is present in vinegar (acetum = vinegar).
IUPAC System: The names of the carboxylic acids are derived from the names of the parent hydrocarbons by replacing the terminal ‘e’ by ‘oic acid’.
Test For Carboxylic acid: 1. When ethanoic acid (acetic acid) is warmed with ethanol in the presence of a few drops of concentrated sulphuric acid, a sweet smelling ester called ethyl ethanoate is formed.
2. Acetic acid produces red colour when a neutral solution of ferric chloride is added to it.
Glacial acetic acid: Acetic acid, when cooled sufficiently it forms ‘ice like’ crystals which melts at 16.7°C. Hence the pure anhydrous acid is usually called glacial acetic acid.
Soda – lime: It is a mixture of caustic soda (NaOH) and quick lime (CaO).

Decarboxylation: The removal of carbon dioxide from a carboxylic acid is known as decarboxylation.

Vinegar: Dilute aqueous solution (5 –8%) of ethanoic acid is called vinegar, which is used to preserve food
Quick vinegar process: Vinegar is prepared by the fermentation of ethyl alcohol with the bacteria acetobacter in the presence of air.
Pyroligneous acid: It is a mixture containing 10% acetic acid, 4% methyl alcohol and 0.5% acetone.

SELF EVALUATION (T.B.Page 155 & 156)

I. Choose the correct answer.

1. The IUPAC name of HCHO is ....
   (a) Methanal (b) Methanol
   (c) Formaldehyde (d) acetaldehyde
2. Formic acid is present in ....
   (a) Ants (b) Vinegar
   (c) Butter (d) none
3. .... percentage composition of acetic acid is called vinegar.
   (a) 10% (b) 0.5%
   (c) 4% (d) 3 – 7 %
4. Pyroligeneous acid is a mixture of ....
   (a) Acetic acid + Methanol + vinegar
   (b) Wood tar + methanol + wood coal
   (c) Acetic acid +methanol + acetone
   (d) Acetic acid + methanol
5. Pure anhydrous acetic acid is known as ....
   (a) Glacial acetic acid (b) vinegar
   (c) acetic acid (d) None
6. The existence of dimer in acetic acid is due to ....
   (a) Intermolecular hydrogen bonding (b) intramolecular hydrogen bonding
   (c) Both a & b (d) None
7. .... is used as a coagulant for rubber latex.
   (a) Ethanoic acid (b)Methanoic acid
   (c) Formaldehyde (d) acetaldehyde
8. .... is a dilute solution of ethanoic acid.
   (a) Vinegar (b) Formic acid
   (c) White lead (d) Paris green
9. Ethanoic acid produces .... colour when treated with neutral ferric chloride.
   (a) Red (b) Violet
   (c) Blue (d) None
10. .... is used in making photographic films.
    (a) Vinegar (b) Cellulose ethanoate
    (c) Paris Green (d) None

Answers:

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

II. Answer the following in One or Two sentences. (T.B.Page 156)

1.What are mixed Ketone? Give an example.

In ketones the two alkyl groups (R & R′) are different, the ketone is said to be a **mixed ketone.**

Example: $CH_3 - CO - CH_2 - CH_3$ - Ethyl methyl ketone or butanone.

2.What are carboxylic acids? Give an example.

Organic compounds which contain the carboxyl functional group (– COOH) **are called** the carboxylic acids.

Example*:* Formic acid - HCOOH; Acetic Acid - CH_3COOH.

3.Give Common Name and IUPAC name for CH_3 COOH?

Common name : Acetic acid

IUPAC name : Ethanoic acid

4.Mention the products obtained on destructive distillation of wood.

When wood is destructively distilled in cast iron retorts, a number of products are obtained. The important ones are-

(i) wood gas

(ii) an aqueous distillate called pyroligneous acid (contains 10% of acetic acid, 4% of methyl alcohol, 0.5% of acetone)
(pyro = fire, lignum = wood)

(iii) wood tar and *(iv)* wood charcoal.

5.CH_3COOH + NaOH $\longrightarrow$ Complete the reaction.

$$\underset{\text{Sodium acetate}}{CH_3COONa} + NaOH \xrightarrow[\Delta]{} \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

III. Answer in brief. (T.B. Page 156 –157)

1.Give two tests for acetic acid.

1. When ethanoic acid (acetic acid) is warmed with ethanol in the presence of a few drops of concentrated sulphuric acid, a sweet smelling ester called ethyl ethanoate (or) ethylacetate is formed.

$$CH_3COOH + C_2H_5OH \longrightarrow \underset{\text{Fruity smell}}{CH_3COOC_2H_5} + H_2O$$

2. Acetic acid produces red colour when a neutral solution of ferric chloride is added to it.

2.What is decarboxylation? Give an example.

Decarboxylation is the elimination of CO_2 from a carboxylic acid. When sodium salts of acetic acid is heated with soda lime (NaOH +CaO) alkanes are formed.

$$\underset{\text{Sodium acetate}}{CH_3COONa} + NaOH \xrightarrow[\Delta]{} \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

3.CH_3 COOH + $NH_3 \longrightarrow$ .....?..... $\longrightarrow$ .....?..... Complete the reaction.

$$\underset{\text{Acetic acid}}{CH_3COOH + NH_3} \longrightarrow \underset{\text{Ammonium acetate}}{CH_3COONH_4} \longrightarrow \underset{\text{Acetamide}}{CH_3CONH_2} + H_2O$$

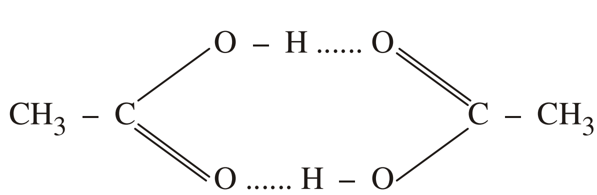
4.Write a note on Quick Vinegar Process.

A dilute solution of acetic acid (3 – 7%) is known as vinegar. It is prepared by the fermentation of ethyl alcohol with the bacteria acetobacter in presence of air.

$$CH_3 - CH_2OH + O_2 \longrightarrow CH_3 - COOH + H_2O$$

5.Acetic acid exists as dimer – Explain.

Acetic acid exists as a dimmer due to intermolecular hydrogen bonding between two molecules.



IV. Answer in detail. (T.B. Page 157)

1.Give common and IUPAC name for the following compounds.

(a) $CH_3 - CO - CH_2 - CH_3$ **(c) $CH_3 - CH_2 - COOH$**

(b) $CH_3 - CH_2 - CO - CH_2 - CH_3$ **(d) HCOOH**

Formula	Common name	IUPAC name
(a) $CH_3 - CO - CH_2 - CH_3$	Ethyl methyl ketone	Butanone
(b) $CH_3 - CH_2 - CO - CH_2 - CH_3$	Diethyl keton	Pentan-3-one
(c) $CH_3 - CH_2 - COOH$	Propionic acid	Propanoic acid
(d) HCOOH	Formic acid	Methanoic acid

2.Discuss the manufacture of acetic acid from pyroligeneous acid.

When wood is destructively distilled in cast iron retorts, a number of products are obtained. The important ones are-

(i) wood gas

(ii) an aqueous distillate called pyroligneous acid (contains 10% of acetic acid, 4% of methyl alcohol, 0.5% of acetone)
(pyro = fire, lignum = wood)

(iii) wood tar and *(iv)* wood charcoal.

Pyroligneous acid solution contains acetic acid (10%), methyl alcohol (4%) and acetone (0.5%). The vapours of pyroligneous acid are passed through hot milk of lime. Acetic acid reacts with milk of lime forming calcium acetate. Methyl alcohol, acetone and water vapours pass off. The solution is evaporated to get dry calcium acetate crystals. These crystals are then treated with concentrated sulphuric acid to obtain acetic acid.

$$(CH_3COO)_2Ca + H_2SO_4 \longrightarrow CaSO_4\downarrow + 2CH_3COOH$$

3.Give any five uses of Ethanoic acid.

(i) as a coagulant for rubber latex

(ii) in the manufacture of plastics, rayon, drugs and silk

(iii) Dilute aqueous solution (3 –7%) of ethanoic acid is called vinegar, which is used to preserve food

(iv) Pure ethanoic acid is used as a solvent and chemical reagent.

(v) As cellulose ethanoate, it is used in making photographic films and rayon.

4.Give any four chemical properties of acetic acid?

(i) Salt formation: Acetic acid reacts with alkalies to form corresponding salts. Acetic acid can donate a proton and form salts with bases *i.e* This shows the acidic nature of acetic acid.

Eg. $\underset{\text{Acetic acid}}{CH_3COOH} + NaOH \longrightarrow \underset{\text{Sodium acetate}}{CH_3COONa} + H_2O$

(ii) Formation of acid halides: Acetic acid react with phosphoruspentahalide to form acid halides (or with thionyl chloride, $SOCl_2$, to form acetyl chlorides).

Eg. $\underset{\text{Acetic acid}}{CH_3COOH} + PCl_5 \longrightarrow \underset{\text{Acetyl chloride}}{CH_3COCl} + POCl_3 + HCl$

$\underset{\text{Acetic acid}}{CH_3COOH} + SOCl_2 \longrightarrow \underset{\text{Acetyl chloride}}{CH_3COCl} + SO_2 + HCl$

(iii) Formation of amides: Acetic acid reacts with ammonia to give salts, which on heating yield acetamide.

Eg. $\underset{\text{Acetic acid}}{CH_3COOH} + NH_3 \longrightarrow \underset{\text{Ammonium acetate}}{CH_3COONH_4} \quad \underset{\text{Acetamide}}{CH_3CONH_2} + H_2O$

(iv) Formation of esters: Acetic acid reacts with ethyl alcohol in the presence of a strong acid catalyst like H_2SO_4 to form esters. The reaction is reversible and the forward reaction is called esterification.

Eg. $\underset{\text{Acetic acid}}{CH_3COOH} + \underset{\text{Ethyl alcohol}}{C_2H_5OH} \overset{H^+}{\quad} \underset{\text{Ethyl acetate}}{CH_3COOC_2H_5} + H_2O$

OTHER IMPORTANT QUESTIONS & ANSWERS

I. Choose the correct answer:

1.The functional group present in aldehyde is
(a) – COOH (b) – CHO (c) – COOR (d) RCOR
2. The IUPAC name of H – CHO is
(a) Methanal (b) Ethanal (c) Propanal (d) Butanal
3. In IUPAC system aldehydes are called
(a) alkanes (b) alkenes (c) alkanals (d) alkynes
4. $RCOR^1$, if the two alkyl groups (R &R^1) different, the compound is said to be
(a) simple ketone (b) mixed ketone (c) normal ketone (d) none
5.The structure of Propanal is
(a) CH_3- CHO (b) H –CHO (c) CH_3 –CH_2 –CHO (d) CH_3 –CO - CH_3
6.The common name for CH_3 –CO –CH_2 –CH_3 is
(a) dimethyl ketone (b) diethyl ketone (c) methyl ethyl ketone (d) ethyl methyl ketone
7.Acetone is a
(a) aldehyde (b) simple ketone (c) carboxylic acid (d) mixed ketone
8. The IUPAC name of methyl n-propyl ketone is
(a) Pentan –3-one (b) Pentan –2-on (c) Pentan –1-on (d) Pentanone
9. The structure of 3 –Pentanone is
(a) $CH_3CH_2COCH_2CH_3$ (b) $CH_3COCH_2CH_2CH_3$
(c) $CH_3CH_2CH_2CH_2CHO$ (d) $CH_3COCH_2CH_3$
10.Organic compound which contain the carboxyl functional group is represented by
(a) R– COOH (b) R– CHO (c) R– COOR (d) RCOR
11.The acid obtained by hydrolysis of vegetable oil is called as
(a) simple acids (b) dicarboxylic acids (c) mono carboxylic acids (d) fatty acids
12.Which one of the following is an example for fatty acids
(a) Palmitic acid (b) acetic acid (c) butyric acid (d) formic acid
13.The acid that present in ants
(a) Ethanoic acid (b) acetic acid (c) Propionic acid (d) Methanoic acid
14.The butterfat contains ………………… acid
(a) formic acid (b) butyric acid (c) acetic acid (d) fatty acid
15.The structure of Propanoic acid is
(a) CH_3CH_2COOH (b) $CH_3CH_2CH_2COOH$ (c) HCOOH (d) CH_3COOH
16.In quick vinegar process ……………. is obtained
(a) Palmitic acid (b) acetic acid (c) butyric acid (d) formic acid
17.Pyroligneous acid solution contains
(a) 20% acetic acid (b) 30% acetic acid (c) 10% acetic acid (d) 100% acetic acid
18. Acetic acid reacts with milk of lime forming …………. .
(a) sodium acetate (b) calcium acetate (c) potassium acetate (d) none
19.The acetic acid exists as a dimmer due to ……………..
(a) inter molecular hydrogen bonding (b) intra molecular hydrogen bonding
(c) Covalent bonding (d) inter molecular carbon bonding
20.The pure 100% acetic acid is called as
(a) anhydrous acetic acid (b) ice like acid (c) glacial acetic acid (d) all of the above
21. Acetic acid on reacts with alkalis to form
(a) sodium acetate (b) potassium acetate (c) calcium acetate (d) all of the above
22. Acid chlorides obtained when acid is react with
(a) PCl_5 (b) $SOCl_2$ (c) both a and b (d) $POCl_3$
23. Acetamide is produced when acetic acid is reacts with.
(a) Ammonia (b) alkalis (c) amides (d) none
24. Soda lime is the mixture of
(a) KOH + CaO (b) NaOH + CaO (c) NaOH + $CaCO_3$ (d) NaOH + Ca $(OH)_2$
25. When carboxylic acids react with alcohols in presents of sulphuric acid to give
(a) Ester (b) ether (c) amine (d) amide
26. Neutral ferric chloride solution gives ………………… colour with acetic acid
(a) Blue (b) Violet (c) Red (d) Black
27. The vinegar solution contains dilute aqueous solution…………….. %
(a) 4 –7% (b) 3 –7% (c) 3 – 4% (d) 7 –9%
28. The IUPAC name of Butaraldehyde is

(a) Butanal (b) Propanal (c) But-1-enal (d) None of these

29.IUPAC name of Acetaldehyde is

(a) butanal (b) 3 butanal (c) Ethanal (d) panal

30. The IUPAC name of the acid which was originally obtained by the distillation of red ants is

(a) Acetic acid (b) Palmitic acid (c) methanoic acid (d) Ethanoic acid

31.Which of the following is not prepared from acetic acids?

(a) Acetyl chloride (b) acetamide (c) ethyl acetate (d) Alkanals

32.Which of the following will not be able to produce acetyl chloride by its reaction with acetic acid?

(a) PCl5 (b) $SOCl_2$ / P_y (c) Cl_2 (d) None of these

33. acetic acid reacts with ……………. to give acetamide

(a) NaOH (b) NH_3 (c) NaOH / Br_2 (d) N_2

34. In esterification reaction, the role of conc. H_2SO_4 is

(a) to act as dehydrating agent (b) to act as hydrolytic agent

(c) to act as catalyst (d) to act as dehydrogenating agent

35. Which of the following is present in Vinegar?

(a) Acetic acid (b) Citric acid (c) Tartaric acid (d) Lactic acid

36.The higher boiling points of Acetic acids are due to?

(a) high density (b) Their trimerisation

(c) Their ability to form intermolecular hydrogen (d) Both (a) and (b)

37. Acetic acid exists as dimmer in benzene due to?

(a) Condensation reaction (b) Presence of –COOH group

(c) Presence of hydrogen atom at β-carbon (d) inter molecular hydrogen bonding

38. Vinegar contains?

(a) Acetic acid (b) Formic acid (c) Sulphuric acid (d) Melonic acid

39.Acetic acid is obtained when?

(a) Methyl alcohol is oxidized with $KmnO_4$

(b) Calcium acetate is distilled in the presence of H_2SO_4

(c) Acetaldehyde is oxidized with $K_2Cr_2O_7$ and HC*l* (d) Glycerol is heated with H_2SO_4

40.Which of the following will give acetamide on reaction with NH3?

(a) CH_3COOH (b) CH_3COCH_3 (c) $(CH_3CO)_2CHOH$ (d) CH_3COOCH_3

41.The IUPAC name of the compound $CH_3 - CH_2 - CHO$ is

(a) propane (b) butanone (c) Butanal (d) propanal

42. When a carboxylic acid is treated with alcohol in presence of a mineral acid the product is

(a) aldehyde (b) ketone (c) ester (d) ether

43. Carboxylic acid group can be detected by

(a) sodium bisulphate test (b) Fehling solution test

(c) Tollen's test (d) Ferricchloride test

44. In the reaction,

$CH_3COOH \xrightarrow{NH_3} A$

(a) Ammonium acetate (b) Acetonitrile (c) Acetic anhydride (d) Ethyl acetate

45. Decarboxylation of sodium Acetate on heating with soda lime gives

(a) Methane (b) Toluene (c) Autaldehyde (d) Acetic acid

46.Which acid is present in vinegar?

(a) Citric acid (b) Tartaric acid (c) Acetic acid (d) Formic acid

47.Calcium acetate on distillation gives

(a) Acetone (b) Acetic acid (c) Acetaldehyde (d) Formicaldehyde

48.Which of the following acid occurs in ants?

(a) Formic acid (b) Acetic acid (c) Propionic acid (d) Oxalic acid

49. The reaction of acids with alcohols is called

(a) Esterification (b) Saponification (c) Hydrolysis (d) Neutralisation

50. $CH_3COOH \xrightarrow{A} CH_3\ COCl$. A is …………………

(a) HCl (b) Cl_2 (c) Cl (d) HCl + Cl_2

Answers:

1.(b) 2.(a) 3.(c) 4.(b) 5.(c) 6.(d) 7.(b) 8.(b) 9.(a) 10.(a) 11.(d) 12.(a) 13.(d) 14.(b) 15.(a) 16.(b) 17.(c) 18.(b) 19.(a) 20.(d) 21.(d) 22.(c) 23.(a) 24.(b) 25.(a) 26.(c) 27.(b) 28.(a) 29.(c) 30.(a) 31.(d) 32.(c) 33.(b) 34.(a) 35.(a) 36.(c) 37.(d) 38.(a) 39.(b) 40.(a) 41.(d) 42.(c) 43.(d) 44.(a) 45.(a) 46.(c) 47.(b) 48.(a) 49.(a) 50.(b)

II. Answer in one or two sentences:

1. Write the IUPAC name of HCHO and CH_3CHO.

HCHO - Methanal

CH_3CHO - Ethanal

2. Give two example of mixed ketone.

1) Butanone - $CH_3–CH_2–CO–CH_3$

2) 2-Pentanone - $CH_3–CO–CH_2–CH_2–CH_3$

3. Give the general formula of aldehydes and ketones?

The general formula of aldehydes is RCHO.

The general formula of ketones is R-CO-R'.

4. Give the structure of propanal and propanone.

1) Propanal - $CH_3–CH_2–CHO$.

2) Propanone - CH_3COCH_3

5. Give the structure and write the IUPAC name of diethyl ketone.

$$\overset{5}{CH_3}\text{-}\overset{4}{CH_2}–\overset{3}{CO}–\overset{2}{CH_2}–\overset{1}{CH_3}$$

IUPAC Name: 3 – pentanone.

6. Give examples for two long chain fatty acids.

Two long chain fatty acids are stearic acid and palmitic acid.

7. What is the origin for the name formic acid and acetic acid?

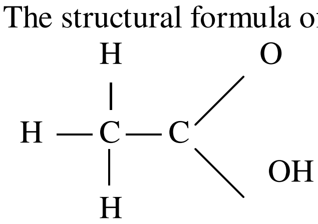
1) Formic acid is present in ants (Latin formica = ants)

2) Acetic acid is present in vinegar (acetum = vinegar)

8. What is glacial acetic acid? Draw the structure.

100% acetic acid which freezes as ice on cooling is called glacial acetic acid.

The structural formula of acetic acid



9. Draw the structure of butanoic acid and propionic acid.

1) $CH_3–CH_2\ CH_2–COOH$ – Butanoic acid.

2) $CH_3–CH_2–COOH$ - Propionic acid.

10. How do you classify carboxylic acids?

Carboxylic acids are classified as monocarboxylic acids, dicarboxylic acids, tricarboxylic acids etc., according to the number of – COOH groups present in the molecule 1, 2, 3, *etc.* respectively.

11. What are called fatty acids? Give an example.

The long chain monocarboxylic acids are commonly called **fatty acids** because many of them are obtained by the hydrolysis of animal fats or vegetable oils. *Eg.* Stearic acid, Palmitic acid.

12. How the common names of carboxylic acids are derived from the original sources?

The common names of carboxylic acid are usually derived from the Latin or Greek word that indicates the original source of the acid. *For example:*

1. Formic acid is present in ants (*Latin:* formica = ants) and acetic acid is present in vinegar (acetum = vinegar).
2. Butyric acid was named from the fact that butter fat (*Latin:* butyrum = Butter) is a significant source of this acid.

13. How will you convert acetic acid into ethylacetate?

Acetic acid reacts with ethyl alcohol in the presence of a strong acid catalyst like H_2SO_4 to form esters. The reaction is reversible and the forward reaction is called esterification.

$$\textit{Eg.}\ \underset{\text{Acetic acid}}{CH_3COOH} + \underset{\text{Ethyl alcohol}}{C_2H_5OH} \xrightarrow{H^+} \underset{\text{Ethyl acetate}}{CH_3COOC_2H_5} + H_2O$$

14.What is called vinegar?

Dilute aqueous solution (3 –7%) of ethanoic acid is called vinegar, which is used to preserve food

III. Answer in brief:

1. How do you name aldehydes by IUPAC system?

In the IUPAC system, the ending of the name (suffix) of aldehyde is ‘al’. The names of aliphatic aldehydes are derived from the name of the corresponding alkane by replacing the terminal ‘e’ by the suffix ‘al’.

(e.g.,) H –CHO – methanal.

2. How do you name aldehydes by trivial system?

In the trivial or common system, aldehydes are named after the carboxylic acids they form on oxidation. The name is obtained by replacing the terminal ‘ic acid’ in the name of the acid by the suffix aldehyde.

(e.g.,) CH_3 –CHO – acetaldehyde.

3. What are the rules for naming ketones in the IUPAC system?

Rules for naming ketones in the IUPAC system:

1. The longest chain containing the CO group is selected and the name of the parent hydro carbon is decided.

2. The last ‘e’ of the name of the hydrocarbon is replaced by ‘one’.

3. The numbering of C atoms in the chain is started from the end which is nearest to the carbonyl group (C=O).

4. How do you name ketones in common system?

In the common system, the names of ketones are obtained by naming the two alkyl groups attached to the keto group (alphabetically) and adding the suffix ‘ketone’.

(e.g.,) CH_3 –CO –CH_3 – Acetone.

5. How do you name ketones in IUPAC system?

In the IUPAC system, the characteristic ending for ketones is ‘-one’. The names of individual aliphatic ketones are derived by replacing the terminal ‘e’ in the name of the corresponding alkane by the suffix –‘one’. (e.g.,) CH_3 –CO –CH_3 – Acetone.

6. How do you name carboxylic acids using trivial name.

The simple carboxylic acids are given common names which are derived from the Latin or Greek word that indicate the original source of the acid.

(e.g.,) (i) Formic acid is present in ants (Latin : formica = ants) and acetic acid is present in vineger (acetum= vinegar)

(ii) Butyric acid was named from the fact that butter fat (Latin = butyrum = butter) is a significant source of this acid.

7. How do you name carboxylic acids using IUPAC system?

IUPAC system:

1. The longest carbon chain containing te carboxyl group is consideed as te parent hydrocarbon.

2. The names of the carboxylic acids are derived from the names of the parent hydrocarbons by replacing the terminal’e’ by ‘oic acid’.

IV. Answer in detail:

1.Explain the naming of ketones in IUPAC system by giving the various rules.

In the **IUPAC system**, the characteristic ending for ketones is -‘one’.

The names of individual aliphatic ketones are derived by replacing the **terminal ‘*e*’** in the name of the corresponding alkane by the suffix -‘one’.

i.e., Name of ketone = Name of corresponding alkane – *e* + one.

Rules for naming Ketones in the IUPAC System

1. The longest chain containing the CO group is selected and the name of the parent hydrocarbon is decided.
2. The last '*e*' of the name of the hydrocarbon is replaced by 'one'
3. The numbering of C atoms in the chain is started from the end which is nearest to the carbonyl group(C = O)

Example 1:

In the compound, $\overset{1}{CH_3} - \overset{2}{CO} - \overset{3}{CH_2} - \overset{4}{CH_2} - \overset{5}{CH_3}$ there are five C atoms, so the parent hydrocarbon is pentane and the name of the ketone is 2-pentanone or pentan-2-one (the CO group gets number 2)

Example 2:

In the compound $\overset{1}{CH_3} - \overset{2}{CH_2} - \overset{3}{CO} - \overset{4}{CH_2} - \overset{5}{CH_3}$ there are five C atoms, so the parent hydrocarbon is pentane and the name of the ketone is 3-pentanone or pentan-3-one (the CO group gets number 3)

2. Explain the naming of aldehydes in common and IUPAC system by giving examples.

In the trivial or common system, aldehydes are named after the carboxylic acids they form on oxidation. The name is obtained by replacing the **terminal 'ic acid'** in the name of the acid **by the suffix aldehyde**.

e.g., HCHO **(or)** $H - \overset{\overset{H}{|}}{C} = O$ Formaldehyde (corresponding acid is formic acid)

$CH_3 - CHO$ **(or)** $CH_3 - \overset{\overset{H}{|}}{C} = O$ Acetaldehyde (corresponding acid is acetic acid)

In the IUPAC system, the ending of the name (suffix) of aldehyde is 'al'. The names of aliphatic aldehydes are derived from the name of the corresponding alkane by replacing the terminal '*e*' by the suffix 'al'.

i.e., Name of aldehyde = Name of corresponding alkane – *e* + al.

Eg: (1) H – CHO

In the above example, one carbon atom is present i.e. the parent hydrocarbon is methane.

$$\text{Methane} \xrightarrow[+al]{} \text{Methanal}$$

(2) CH_3 – CHO (2 -carbon atoms present)

$$\text{Ethane} \xrightarrow[+al]{} \text{Ethanal}$$

(3) $CH_3 - CH_2 - CHO$ (Propanal)

The common and IUPAC names of some aldehydes are given below:

Formula	Common name	IUPAC name
H – CHO	Formaldehyde	Methanal
$CH_3 - CHO$	Acetaldehyde	Ethanal
$CH_3 - CH_2 - CHO$	Propionaldehyde	Propanal
$CH_3 - CH_2 - CH_2 - CHO$	Butyraldehyde	Butanal

3. How do you name monocarboxylic acid in common and IUPAC system? Explain with examples.

The simple carboxylic acids are better known by their common names. The common names are usually derived from the Latin or Greek word that indicates the original source of the acid. For example:

a. Formic acid is present in ants (*Latin:* formica = ants) and acetic acid is present in vinegar (acetum = vinegar).

b. Butyric acid was named from the fact that butter fat (*Latin:* butyrum = Butter) is a significant source of this acid.

IUPAC system

1. The longest carbon chain containing the carboxyl group is considered as the parent hydrocarbon.
2. **The names of the carboxylic acids are derived from the names of the parent hydrocarbons by replacing the terminal 'e' by 'oic acid'.**

i.e., **Name of carboxylic acid = Name of corresponding alkane –e + oic acid**

Example: (1) HCOOH

$$\text{Methane} \xrightarrow[+oic\ acid]{} \text{Methanoic acid}$$

(2) CH_3COOH

$$\text{Ethane} \xrightarrow[+oic\ acid]{} \text{Ethanoic acid}$$

The common and IUPAC names of some of the acids are given below:

Formula	Common name	IUPAC name
$HCOOH$	Formic acid	Methanoic acid
CH_3COOH	Acetic acid	Ethanoic acid
CH_3CH_2COOH	Propionic acid	Propanoic acid
$CH_3CH_2CH_2COOH$	Butyric acid	Butanoic acid

4. Write a short note on physical properties of acetic acid.

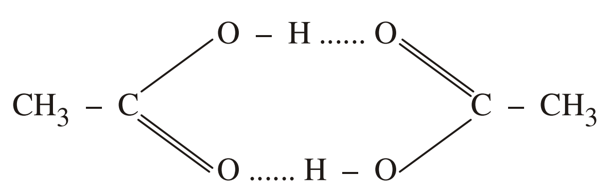
(i) Acetic acid is a colourless liquid (boiling point 118°C) with a sharp 'vinegar odour' and sour taste.

(ii) It is miscible with water, ethyl alcohol and ether in all proportions.

(iii) When cooled sufficiently, it forms **'ice like'** crystals which melt at 16.7°C. Hence the pure anhydrous acid is usually called **glacial acetic acid** (Glacial = of ice).

(iv) Acetic acid exists as a dimmer due to intermolecular hydrogen bonding between two molecules.

(v) It is a good solvent for sulphur, phosphorus, iodine and many organic compounds.



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11 – CHEMISTRY AND ENVIRONMENT

INTRODUCTION

Environment constitutes air, water, soil, the atmosphere and the plants that are around us. Environmental chemistry deals with the chemical phenomena taking place around us, and the impact it has on the environment. It is also related to biology, agriculture, medicine and public health. In recent years air, water and soil are polluted heavily due to human activity causing threat to the very existence of life on the planet, earth. Earth is the only planet in the solar system having conditions required for the survival of the living organisms. In view of this, it is not only necessary but also essential for various living organisms to live in harmony – a kind of mutual co-existence on earth is the need of the hour. In recent years, the mutual co-existence has been very much disturbed by human activity and the whole world is now trying hard to protect the environment. It has therefore, become necessary to educate people with regard to the environment, the damage caused to it due to the activities of the people, and the consequences. It is for this reason that the subject has been included for study at all levels, for the students. This subject also emphasizes the measures to be taken to protect the environment.

IMPORTANT TERMS & DEFINITIONS

Pollution: It is defined as an addition or excessive addition of undesirable materials to the physical environment (air, water, and land), making it less fit or unfit for life.
Basic Cause of Pollution: - There are two main causes of pollution
(1) Human activities and (2) Natural phenomena.
Pollutant: A substance released into the environment due to natural or human activity and effect adversely the environment is called as pollutant, eg : sulphur dioxide, carbon monoxide, lead, mercury etc.
Receptor: The medium which is effected by the pollutant is called receptor.
Eg: When many vehicles stop at the traffic signal during peak hours, our eyes become red with burning sensation due to the smoke released from the automobiles. The eyes here are the receptors.
Sink: The medium which reacts with pollutants is called sink. Eg: Micro-organisms which eat the dead animals or which convert the dried leaves and garbage into fertilizers. Thus, the pollutant is removed by micro -organisms. Similarly, seawater is a big sink for carbon dioxide.

Classification

Quantitative pollutants

These are the substances, which normally occur in nature but are also added in large qualities by man. For instance: carbon dioxide. It is always present in the air, and is also released by fires, industries and automobiles.

Qualitative pollutants

These are the substances which do not occur in nature but are added by man. The insecticides and herbicides, for example, are qualitative pollutants.

Biodegradable pollutants

These are quickly degraded by natural means. Sewage and heat are pollutants of this category. These pollutants are disposed of by microbial action and radiation.

Non degradable pollutants

These are not degraded or are degraded very slowly in nature. D.D.T., arsenic salts of heavy metals, glass or tin containers, radioactive materials, and plastic are the pollutants of this category. These pollutants accumulate and may get biologically magnified as they pass through the food chains.

Primary pollutants

These persist in the form in which they added to the environment. Plastic ware are primary pollutants.

Secondary pollutants

These are formed by interaction among the primary pollutants. For example, two primary pollutants, namely, nitrogen oxides and hydrocarbons, from motor vehicles, react in the presence of sunlight to form two secondary pollutants, viz., peroxyacyl nitrate (PAN) and ozone. These are more toxic than the primary pollutants. This phenomenon of increased toxicity by reaction among the pollutants is called synergism.

Types of pollution: (i) Air pollution (ii) Water pollution. (iii) Soil or land pollution
Air pollution: It refers to the release into the atmosphere, of materials that are harmful to man, other animals, plants and buildings or other objects.

Causes of air pollution: There are two main causes of air pollution: human activities and natural phenomena.

(a) Human Activities

Water Pollution: It is defined as the addition of some foreign substance (organic, inorganic, biological or radiological) to water, or change in its physical property (heat) that constitutes a health hazard or otherwise make it less fit or unfit for use.

Soil Pollution: Alteration in soil by addition and removal of materials leading to reduce productivity is called soil pollution.

Soil Pollutant: Substances which reduce productivity of the soil are regarded soil pollutants.

Type of soil pollution: Soil Pollution is of two main types: Positive and negative.

Green chemistry: It is defined as the design of chemical products and processes that reduce or eliminate the use and generation of hazardous substances. The ultimate aim of green chemistry or environment friendly chemistry is to prevent pollution at the source**.**

SELF EVALUATION (T.B. Page 173)

I. Choose the correct answer.

1. The medium which is affected by the pollutant is called ....
   (a) Pollutant (b) Pollution
   (c) Sink (d) Receptor
2. D.D.T is an example for .... type of pollutant
   (a) Degradable (b) Bio- Degradable
   (c) Non-Bio Degradable (d) None
3. The main sources of sulphurdioxide pollutant is ....
   (a) Volcanic activity (b) Chemical industry
   (c) Power houses (d) Textile mills
4. Lower concentration of SO_2 in air causes ....
   (a) Shortness of breath (b) Cough
   (c) temporary spasms of larynx (d) cancer
5. Emphysema is caused by the .... pollutant
   (a) CO (b) SO_2
   (c) Hydrocarbons (d) particulate
6. Poisonous gases can be removed by ....
   (a) the spray collector (b) afforestation
   (c) Filteration (d) using a catalyst
7. The presence of cyanides in water causes ....
   (a) Change in acidity or alkalinity of H_2O
   (b) Decrease in photosynthesis
   (c) Eutrophication
   (d) none
8. Environment friendly chemistry is otherwise is known as ....
   (a) Green chemistry (b) Industrial chemistry
   (c) Grganic chemistry (d) none
9. Pyrolysis is ....
   (a) Burning with oxygen (b) burning without oxygen
   (c) both a & b (d) none of the above
10. Asbestosis is the disease which affect ....
   (a) Lungs (b) heart
   (c) eye (d) blood

Answers:

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

II. Answer the following in One or Two sentences.(T.B. Page 174)

1.Define Pollution.

Definition 1: Pollution may be defined as a change in the physical, chemical or biological aspects of the environment which make it harmful for humans, other living organisms and cultural assets.

Definition 2: Pollution may be defined as an addition or excessive addition of undesirable materials to the physical environment (air, water, and land), making it less fit or unfit for life.

2.What are secondary Pollutants?

These are formed by interaction among the primary pollutants. For example, two primary pollutants, namely, nitrogen oxides and hydrocarbons, from motor vehicles, react in the presence of sunlight to form two secondary pollutants, viz., peroxyacyl nitrate (PAN) and ozone. These are more toxic than the primary pollutants.

3.Define Air pollution.

Air pollution refers to the release into the atmosphere, of materials that are harmful to man, animals, plants and buildings or other objects.

4.What are the changes occurring in H_2O due to pollutants?

If the pollutants are present in water, the following changes occur.

1. Change in the colour and increase in the salinity of water.
2. Bad odour starts emanating from rivers, ponds, and lakes.
3. Uncontrolled growth of the weeds in water.
4. Decrease in the growth of fish.

5.Mention the types of soil pollutants.

The various soil pollutants are:

(i) Chemicals

(ii) Pesticide

(iii) Fertilizers and organic manure,

(iv) Radioactive wastes, and

(v) Discarded materials.

6.Define noise pollutants.

Noise can be defined as an unwanted sound at a wrong place, at the wrong time. Although noise pollution is not caused by chemicals it is equally harmful.

III. Answer in brief. (T.B. Page 174)

1.Mention the measures to avoid pollution.

*(i)*Use of automobiles should be minimised. This will not only reduce pollution of air but will also conserve oil and prove economical.

(ii) Conventional fuels (firewood, coal, oil) should be replaced by electricity or natural gas. These fuels do not emit SO_2.

(iii) Population growth, the main cause of pollution, should be brought under control.

(iv) Nuclear explosions and wars should be stopped.

2.How can you reduce sulphur content in air?

By using low sulphur fuel in motor vehicles we can reduce sulphur content in air.

3.Mention the sources of air pollution.

The major air pollutant are listed below in table:

S.No.	Pollutant	Main sources
1.	Carbon monoxide	Fossil fuel burning, furnaces, power houses
2.	Sulphur oxides	Volcanic activity, refineries.
3.	Nitrogen di-oxides	Chemical industries
4.	CFC (chlorofluorocarbons)	Aerosol propellants, chemicals used for dry cleaning
5.	Dust	Asbestos and cement industries, Textile mills.

4.Define receptor with an example.

The medium which is affected by the pollutant is called receptor. *Eg:* When many vehicles stop at the traffic signal during peak hours, our eyes become red with a burning sensation due to the smoke released from the automobiles. The eyes here are the receptors.

5.Define quantitative pollutants with example.

These are the substances, which normally occur in nature but are also added in large quantities by man. For instance: carbon dioxide. It is always present in the air, and is also released by fires, industries and automobiles.

IV. Answer in detail. (T.B. Page 175)

1.Give a general account on the classification of pollutants.

The pollutants are classified based on the different points, which are discussed below:

(a). According to their existence in nature, pollutants may be quantitative or qualitative.

Quantitative pollutants: These are the substances, which normally occur in nature but are also added in large quantities by man. For instance: carbon dioxide. It is always present in the air, and is also released by fires, industries and automobiles.

Qualitative pollutants: These are the substances which do not occur in nature but are added by man. The insecticides and herbicides, for example, are qualitative pollutants.

(b). According to their natural disposal, the pollutants may be biodegradable or non degradable.

1. Biodegradable pollutants: These are quickly degraded by natural means. Sewage and heat are pollutants of this category. These pollutants are disposed off by microbial action and radiation.

2. Non degradable pollutants: These are not degraded or degrad very slowly in nature. D.D.T., arsenic salts of heavy metals, glass or tin containers, radioactive materials, and plastics are the pollutants of this category. These pollutants accumulate and may get biologically magnified as they pass through the food chains.

(c). According to the form in which they persist after release into the environment, the pollutants may be primary or secondary.

1. Primary pollutants: These persist in the form in which they are added to the environment. Plastics are primary pollutants.

2. Secondary pollutants: These are formed by interaction among the primary pollutants. For example, two primary pollutants, namely, nitrogen oxides and hydrocarbons, from motor vehicles, react in the presence of sunlight to form two secondary pollutants, viz., peroxyacyl nitrate (PAN) and ozone. These are more toxic than the primary pollutants. This phenomenon of increased toxicity by reaction among the pollutants is called **synergism**.

2.What are the causes of air pollution?

There are two main causes of air pollution:

1. Human activities and 2. Natural phenomena.

1. Human Activities: Man has been polluting the air ever since he started using fire. Industrialisation and invention of automobiles have speeded up the pollution of air. Overpopulation, deforestation, nuclear explosions and explosives used in wars, and fireworks on festivals are also contributing to air pollution.

2. Nature:

(i) Volcanic eruptions release gases and ashes which pollute the air. Pollution of air volcanic eruptions in certain geological periods seems to have changed the earth's climate.

(ii) Fumes & Fires release harmful gases.

(iii) Natural organic and inorganic decays release harmful dust and sulphurous gases.

(iv) Dust storms are another factor in the pollution of air.

(v) Pollen, spores, cysts, bacterial and marsh gas are natural pollutants.

(vi) Atmospheric pollution existed even before the evolution of man. Man has only aggravated the air pollution.

3.Discuss the types of soil pollution.

There are two types of soil pollution.

1. Positive Soil Pollution: Reduction in productivity of the soil due to the addition of undesirable substances (industrial wastes, pesticides, inorganic fertilizers, radioactive dust and discarded materials) is called positive soil pollution.

2. Negative Soil Pollution: Fertility of the soil depends on the minerals it contains. Minerals abound in the top layer of the soil. Hence, the top layer must remain intact. The factors that reduce the mineral contents of the top layer or damage the top layer reduce the soil fertility. The loss of soil productivity by reduction in its mineral contents or by destruction of its top layer is termed negative soil pollution.

4.How can we conserve energy?

The use of energy is increasing day by day, not only in industry but also in homes. The day is not far, when we will face problems because of its shortage. Due to limited supply and rising demand, the price of energy sources is steadily increasing. To avoid a situation like energy crisis, we should conserve energy.

Judicious use of available energy can help in overcoming the energy crisis. The following steps can help us save energy. Energy saved is energy produced.

1. Drive a scooter or a car only when it is very essential and there is no other alternative. Use of car by a single person is wasteage of petrol. Use of public transport system for travelling can save petrol.
2. In homes, schools, colleges, offices, industry, hospitals, and other places, lights, fans and coolers should be switched off when no one is inside to use it.
3. Modern smokeless chullahs should be used for cooking, as it avoids wasteage. Open chullahs waste about 90% of the fuel.
4. The use of various devices based on solar energy such as solar water heater, solar cooker etc. Should be encouraged.
5. Whenever possible, the use of renewable source should be preferred than using non-renewable source.
6. In villages bio-gas can be produced easily. So its use should be encouraged. Use of pressure cooker also saves energy and time.
7. Leakage at all levels whether in water pipe, gas pipe or oil pipe should be repaired immediately.

5.What are the effects of air pollution on human beings?

Effects of Air Pollution on Human Beings:

1. Gaseous as well as particulate air pollutants cause severe damage to respiratory system leading to emphysema, bronchitis and asthma.
2. Carbon monoxide when inhaled reacts with haemoglobin in blood and reduces its oxygen carrying capacity. This may cause serious injuries to vital organs.
3. Lower levels of SO_2 cause temporary spasm of bronchial muscles. Higher concentration of SO_2 cause shortness of breath, cough and spasm of larynx.
4. Nitrogen oxides cause pulmonary haemorrahage in higher concentration.
5. Hydrocarbon released from automobile exhaust can cause lung cancer.
6. Particulates such as asbestos cause scarring of lungs known as **asbestosis**. Dust particles from silicon containing rocks cause **silicosis**.
7. Lead from automobile exhaust gets accumulated in the body damages bones.

6.How do you minimise noise pollution? What measures can be taken for an effective control of noise pollution?

It is, however, impossible to have a total elimination of annoying sounds. Noise pollution can however be minimised by taking the following important measures.

(a) Reduction of noise at source.
(b) Reduction of population exposed.
(c) Duration of noise exposure, *etc.*

Following measures may be taken for an effective control on noise pollution.

(a) Using ear protective aids.
(b) Proper designing of doors and windows
(c) Improving in working methods
(d) Tree planting
(e) Providing enclosures
(f) Treatment of walls, floors and ceilings
(g) Use of silencers

OTHER IMPORTANT QUESTIONS & ANSWERS

I. CHOOSE THE CORRECT ANSWER:

1. CFC is related to pollution of
(a) Air (b) Water (c) Soil (d) All the above

2. Most hazardous metal pollutant of automobile exhaust is
(a) Mercury (b) Tin (c) Cadmium (d) Lead

3. DDT is
(a) Biodegradable pollutant (b) Nondegradable contaminent
(c) Air pollutant (d) An antibiotic

4. Water pollution is due to
(a) Agricultural discharges (b) Sewage and other wastes (c) Industrial effluents (d) All the above

5. Which of the following reacts with haemoglobin of blood and produce toxic effect.
(a) Carbon dioxide (b) Carbon monoxide (c) Oxygen (d) Carbon suboxide

6. Which of the following is major sink for carbon monoxide?
(a) Water (b) Soil (c) animal respiration (d) Salts dissolved in ocean water
7. Environmental pollution refers to
(a) peeling of topsoil (b) dissipation of energy
(c) release of toxic / undesirable materials in environment (d) None of the above
8. Which of the following is biodegradable pollutant?
(a) Domestic waste (b) DDT (c) Mercury salts (d) Aluminium foil
9. Chief source of water and soil pollution is
(a) Mining of ores (b) Thermal power plant (c) Agro –industry (d) All the above
10.Which of the following cause water pollution?
(a) Smoke / fly ash (b) Automobile exhausts (c) Aeroplanes (d) Silt and pesticides
11.Air pollution is not caused by
(a) Pollen grains (b) Hydroelectric power (c) Industries (d) Automobiles
12.Carbon monoxide is harmful to human beings as it
(a) is carcinogenic (b) is antagonistic to CO_2
(c) has higher affinity for haemoglobin as compared to oxygen (d) is destructive to O_3
13.Which of the following is atmosphere pollutant
(a) CO_2 (b) CO (c) O_2 (d) N_2
14.Which of the following is not a chemical pollutant?
(a) Carbon dioxide layer (b) Oxygen layer (c) Ozone layer (d) Troposphere
15. Oxides of sulphur and nitrogen are important pollutants of
(a) Water (b) Air (c) Soil (d) All the above
16. Silencers is used for minimising
(a) Radioactive pollution (b) Air pollution (c) Noise pollution (d) water pollution
17. Water pollution is mainly due to which of the following
(a) CO_2, CO and SO_2 (b) Nitrous oxide, nitric oxide and nitric acid
(c) Oxygen, chlorine and nitric acid (d) Ozone, chlorine and sulphur dioxide
18. Spraying of DDT produces pollution of
(a) Air (b) Air and water (c) Air and soil (d) Air, water and soil
19. The major air pollutant is:
(a) CO (b) Oxides of nitrogen (c) Soot (d) oxides of sulphur
20.The major source of CO pollution is
(a) industrial processes (b) Power houses (c) forest fires (d) volcanic activity
21.Which one is not a pollutant normally?
(a) Hydrocarbons (b) Carbon dioxide (c) Carbon monoxide (d) sulphur dioxide
22. Sulphur dioxide cause
(a) Injuries (b) pulmonary heamorrahage (c) cough (d) none
23. Pollutant of automobile (eutrophication). Nervous system / produces mental diseases is
(a) Mercury (b) Lead (c) Nitrogen oxide (d) Sulphur oxide
24. Domestic waste mostly constitutes
(a) Non – biodegradable pollution (b) Biodegradable pollution (c) Effluents (d) air pollution
25. DDT is
(a) Green house gas (b) Degradable pollutant
(c) Nondegradable pollutant (d) None of the above
26. CO emissions can be decreased by
(a) Burning fuels at higher temperature (b) Mixing combustion reactants more thoroughly
(c) Passing combustion products over hot charcoal (d) All the above
27.Pollution is
(a) Removal of top soil (b) Release of toxic / undesirable materials in environment
(c) Conservation of energy (d) All the above
28. Lead is
(a) Air pollutant (b) Soil pollutant (c) Radioactive pollutant (d) Noise pollutant
29.Water pollution is mainly due to
(a) Sulphur dioxide (b) carbon dioxide (c) Oxygen (d) industrial discharges
30. Bio – degradable pollutant is
(a) Domestic waste (b) DDT (c) Mercury salt (d) Aluminium foil
31.The main agent for polluting the environment is
(a) Pig (b) Plant (c) man (d) All the above
32.Radioactive wastes are
(a) Biodegradable materials (b) Non- biodegradable materials
(c)Materials that do not cause pollution (d) all the above

33. Nitrogen oxides an hydrocarbons released b automobiles interact to form
(a) CO (b) SO_2 (c) PAN and Ozone (d) Aerosols
34.Which of the following are the by products of burning of fossil fuels
(a) CO (b) CH_4 (c) SO_2 (d) NO_x
35. Major sources of NO pollution are?
(a) Textile mills (b) power house (c) chemical industries (d) All
36. Which of the following is the source of SO_2 pollution?
(a) Volcanic activity (b) Fossil fuel combustion (c) Thermal power plants (d) All
37. Which of the following is not vehicular pollutant?
(a) CO (b) Particulate matter (c) hydrocarbons (d) All are correct
38. Which of the following is secondary pollutant
(a) Aerolein (b) Formaldehyde (c) PAN (d) All
39.Which part of the body is usually damaged from hydrocarbon poisoning?
(a) Heart (b) Brain (c) Kidney (d) Lungs
40. Which is responsible for damaging blood?
(a) Pb (b) As (c) Ca (d) Mg
41.Biodegradable materials are?
(a) Those which spoil the biological environment (b) Are toxic
(c) can be broken down by bacteria (d) Used for converting waste into greenery
42. The green house gas is?
(a) CO_2 (b) CH_4 (c) N_2O (d) CFCs
43. Ozone depletion in the stratosphere is mainly caused by?
(a) SO_2 (b) NO_2 (c) NO (d) CFCs
44.World environment day is?
(a) 5 May (b) 5 June (c) 5 July (d) 5 August
45. Haemoglobin of the blood forms carboxy haemoglobin with?
(a) CO_2 (b) CO (c) SO_2 (d) NO_2
46. Green house effect causes?
(a) Rise in temperature of the earth (b) Continuous rainfall
(c) Lowering in temperature of the earth (d) Continuous snowing on the earth
47. Green plants during daytime absorb?
(a) O_2 (b) CO_2 (c) CO (d) N_2
48. Which is not a green house gas?
(a) CH_4 (b) CO_2 (c) Chlorofluorocarbons (d) All are correct
49. O_3
(a) Is a mild oxidizing and bleaching agent
(b) forms ozonides with unsaturated organic compounds
(c) Layer in stratosphere shields the earth from solar U.V. radiations
(d) All the above
50. Which is not a primary air pollutants?
(a) Oxides of carbon (b) Oxides of S (c) Oxides of N_2 (d) H_2SO_4

Answers:

1.(a) 2.(d) 3.(b) 4.(d) 5.(b) 6.(d) 7.(c) 8.(a) 9.(c) 10.(d) 11.(a) 12.(c) 13.(b) 14.(d) 15.(d) 16.(c) 17.(c) 18.(d) 19.(a) 20.(b) 21.(b) 22.(c) 23.(b) 24.(a) 25.(c) 26.(b) 27.(b) 28.(a) 29.(d) 30.(a) 31.(d) 32.(b) 33.(b) 34.(b) 35.(c) 36.(a) 37.(c) 38.(c) 39.(d) 40.(a) 41.(a) 42.(b) 43.(d) 44.(a) 45.(b) 46.(a) 47.(b) 48.(d) 49.(c) 50.(c).

II. Answer in one or two sentences:

1. Give the main reasons for pollution?

Increase in the population, urbanization, deforestation, etc., are the main reasons for pollution.

2. Give few examples of pollutants.

Sulphur dioxide, carbon monxide, lead, mercury, etc., are some examples of pollutants.

3. What are biodegradable pollutants? Give examples.

Pollutants which are quickly degraded by natural means are called biodegradable pollutants. E.g., sewage and heat.

4. What are nondegradable pollutants? Give examples.

Pollutants which are not degraded or degrade very slowly in nature are called nondegradable pollutants. E.g., DDT, arsenic salts of heavy metals, glass or tin containers, radioactive materials, plastic, etc.

5. What are primary pollutants? Give examples.

The pollutants which in the form in which they are added to the environment. E.g., plastic.

6. Give the general classifications and source of pollutions.

The classifications of pollutions are (i) Air pollution, (ii) Water Pollution, (iii) Soil or land pollutions. Man is the principal source of pollution.

7. What causes (a) asbestosis (b) silicosis?

Particulates such as asbestos cause scarring of lungs known as asbestosis. Dust particles from silicon containing rocks cause silicosis.

8. Give the effects of water pollution?

The affects of water pollution are

1. Polluted water cannot be used for drinking.
2. Polluted water leads to diseases like cholera, jaundice, typhoid and diarrhoea.
3. Aquatic life gets destroyed.

9.What do you mean by synergism?

These are formed by interaction among the primary pollutants. For example, two primary pollutants, namely, nitrogen oxides and hydrocarbons, from motor vehicles, react in the presence of sunlight to form two secondary pollutants, viz., peroxyacyl nitrate (PAN) and ozone. These are more toxic than the primary pollutants. This phenomenon of increased toxicity by reaction among the pollutants is called **synergism**.

10.Define green chemistry.

Green chemistry is defined as the design of chemical products and processes that reduce or eliminate the use and generation of hazardous substances. The ultimate aim of green chemistry or environment friendly chemistry is to prevent pollution at the source.

11.Define environment.

Environment can be defined as the physical surroundings and conditions, which affect the lives of the people, plants and animals.

12. Give the various components of environment?

Environment is composed of three main components namely, lithosphere (land), hydrosphere (water) and atmosphere (air).

13. Name any two basic causes of pollution.

There are two main causes of pollution

1. Human activities and
2. Natural phenomena.

14. Define pollutant.

A substance released into the environment due to natural or human activity and which affects adversely the environment is called as pollutant, *Eg :* sulphur dioxide, carbon monoxide, lead, mercury etc.

15.Define water pollution.

Water pollution is defined as the addition of some foreign substance (organic,

inorganic, biological or radiological) to water, or change in its physical property (heat) that constitutes a health hazard or otherwise make it less fit or unfit for use.

16. Give the main cause of water pollution?

Man is the main cause of water pollution. Some pollution occurs naturally too. Soil particles enter water by its erosion; minerals dissolve in water from rocks and soil; animal wastes and dead fallen leaves fall into water sources. Decaying of organic matter also pollutes water.

17.What is called soil pollution?

Alteration in soil by addition and removal of materials leading to reduce productivity is called soil pollution. Here, soil productivity includes both the quantity and the quality of the produce.

III. Answer in brief:

1. What can we do to reduce sulphur content in air?

Good quality fuel (low sulphur or sulphur free and lead free fuel) should be used in motor vehicles. The exhaust gases from motor vehicles may be cleaned by using catalyst. These steps can reduce sulphur content in air.

2. Explain - conservation and protection of environment.

Conservation means ‘to keep safe’ whereas preservation means ‘to maintain the environment as it is.’ It is necessary to conserve and protect our environment because the natural resources are getting depleted and environmental problems are increasing day by day.

3. What is called energy crisis?

The use of energy is increasing day by day, not only in industry but also in homes. The day is not far, when we will face problems because of its shortage. Due to limited supply and rising demand, the price of energy sources is steadily increasing. To avoid a situation like energy crisis, we should conserve energy. Judicious use of available energy can help in overcoming the energy crisis. Energy saved is energy produced.

4. What is called sink? Give example.

The medium which reacts with pollutants is called sink. *Eg:* Micro-organisms which eat the dead animals or which convert the dried leaves and garbage into fertilizers. Thus, the pollutant is removed by micro - organisms. Similarly, seawater is a big sink for carbon dioxide.

5. Pollution has become a “global phenomenon” – Explain.

Due to increase in the population and industrialisation, the natural resources have diminished. Man has been trying to prepare many natural things artificially and in this process many industries were started. New technologies were introduced in industries for improving the yields. Along with this development, many waste products produced were released into the environment, polluting it badly. This is the case not only with developed countries but also with under developed countries and so pollution has become a **“global phenomenon”**.

Increase in the population, urbanization, deforestation *etc.,* are the main reasons for pollution. With deforestation, many of the wild animals and birds are also getting extinct. Water of the rivers, ponds etc. are getting polluted. As some of the pollutants are entering inner layers of the earth, fertility of the soil is also getting affected and lands are becoming less useful for cultivation. Thus, life on earth is in great danger.

6. Give your suggestions to conserve and protect our environment?

(i) The practice of crop rotation helps in conserving soil.

(ii) Judicious use of fertilisers, intensive cropping, proper irrigation and drainage help in the conservation of soil.

(iii) The treatment of sewage prevents pollution of water bodies and helps in conserving fish and other aquatic life.

(iv) Natural parks and wildlife sanctuaries should be established throughout the country in order to protect and conserve wild animals, birds and plant species.

(v) New trees should be planted in place of those cut for various purposes, which will protect the earth from excessive heating.

(vi) Harvesting of rainwater helps in the conservation of groundwater. Composting of solid organic waste for biogas and manure.

IV. Answer in detail:

1. Discuss the various methods followed to control soil and landscape pollution?

Control of soil and Landscape pollution mainly involves the disposal of solid wastes. Burning of solid wastes pollutes air besides leaving a large amount of residue for further disposal. Pyrolysis (burning without oxygen) is very costly and consumes energy in large amounts. Recycling of wastes is practicable and also beneficial. It not only reduces the amount of residue but also conserves the natural resources.

1. Recycling of wastes: Some important ways of recycling of wastes are given below.

(i) Agricultural Wastes: Agricultural wastes including paddy husk, corncobs, remains of crushed sugarcane, fibrous coat of coconuts, tobacco waste, cereal stems and others are converted into paper and board. Recycling of paper is costly but worthwhile in view of conservation of resources. It is estimated that recovery of one tonne of paper can save 17 trees.

(ii) Waste Paper: Old books, newspapers, magazines, notebooks, answer books, *etc.,* are converted into new paper by paper mills.

(iii) Jute: Jute waste is changed into hardboard.

(iv) Cattle -dung: Cattle dung used now a days in gobar gas plants provides cooking gas and enriched dung manure.

(v) Composting: Domestic wastes (fruit and vegetable peels, fallen leaves) can be disposed of in one's own house by composting. This practice not only reduces environmental pollution but also provides humans to replenish depleted soil resources.

(vi) Water: Clean water resulting from treatment of sewage and industrial wastes can be reused.

2.How will you control air pollution by separating the pollutants? Explain.

Separation of Pollutants: This can be done by following steps

(i) Trees should be grown in all available places. The trees use carbon dioxide and release oxygen. This purifies the air for man and animal to breathe.

Certain plants (phaseolus vulgaris, coleus blumeri) can fix carbon monoxide, and some plants (pinus, juniperus) can metabolise nitrogen oxides. Plantation of such species should be encouraged to depollute the air.

(ii) Good quality fuel (low- sulphur or sulphur- free and lead-free fuel) should be used in motor vehicles. The exhaust gases from motor vehicles may be cleaned by using catalyst. These steps can reduce SO_2 pollution.

(iii) The use of tall chimneys in factories can reduce, pollution of air at ground level.

(iv) Industrial smoke should be filtered before releasing it into the air to remove particulate matter.

(v) Poisonous gases should be removed by passing the fumes through water tower scrubber or spray collector.

(vi) Mining area should be afforestated.

The above said steps may be followed to control air pollution.

3.How will you control air pollution by avoiding the pollutants? Explain.

Avoidance of Pollutants: This can be done by following measures:

(i) Use of automobiles should be minimised. This will not only reduce pollution of air but will also conserve oil and prove economical.

(ii) Conventional fuels (firewood, coal, oil) should be replaced by electricity or natural gas. These fuels do not emit SO_2.

(iii) Population growth, the main cause of pollution, should be brought under control.

(iv) Nuclear explosions and wars should be stopped.

The above said steps may be followed to control air pollution.

4. Discuss the various types of water pollutants and their evil effect.

No.	Class of pollutant	Effect
1.	Salts, trace elements like copper, zinc, arsenic *etc.,* metals coming out from chromium plating industry.	Affects the human health and aquatic animals.
2.	Metals and complex compounds.	Metals disturb the water system. Algae cannot grow properly. Such surroundings decrease photosynthesis and increase air pollution directly.
3.	Cyanides, hydrogen sulphides, carbon dioxide, nitrogen dioxide and sulphites.	Acidity or alkalinity of the water changes and becomes toxic to aquatic animals and plants.
4.	Nutrients like carbon dioxide, hydrogen, oxygen, nitrogen, nitrates, phosphates, sulphates and micro nutrients like boron, chlorine, copper, iron, manganese, vanadium, zinc etc., compounds.	Eutrophication of the pond causes excess growth of the algae and subsequently the ponds get dried up.

5. Discuss the various measures followed to control water pollution?

Pollution of water can be checked, or at least minimised, by the following measures -

1. Taking bath and washing clothes directly in ponds, tanks and streams, which supply drinking water for humans, should be prohibited.

2. Separate ponds and tanks should be reserved for the water supply to cattle and other animals.
3. Domestic and farmyard sewage and industrial waste should be suitably “treated” before releasing them into water. This process can reduce the harmful effect of the wastes.
4. Over use of fertilizers and pesticides should be avoided. As far as possible less stable pesticides should be used.
5. Hot water should be cooled before release from factories. Solid wastes should be recycled wherever possible.

6. Discuss the various evil effects of soil pollutants in detail.

(i) The Chemicals and pesticides alter the basic composition of the soil. This may kill essential soil organisms which contribute to the structure and fertility of soil.

(ii) The chemicals and pesticides may also make the soil toxic for plant growth. Many pesticides or their degradation products are absorbed by plants and may reach animals and humans via food chains, and prove harmful. The use of inorganic fertilizers spoils the quality of the soil in the long run.

(iii) Use of human and animal excreta as manure pollutes the soil besides promoting crop yield. Excreta may contain pathogens that contaminate the soil and vegetable crops and affect the health of man and domestic animals.

(iv) However, biological pollutants play only a minor role in changing soil composition. Radioactive dust may find its way from the soil into crops, livestock and humans via food chains.
