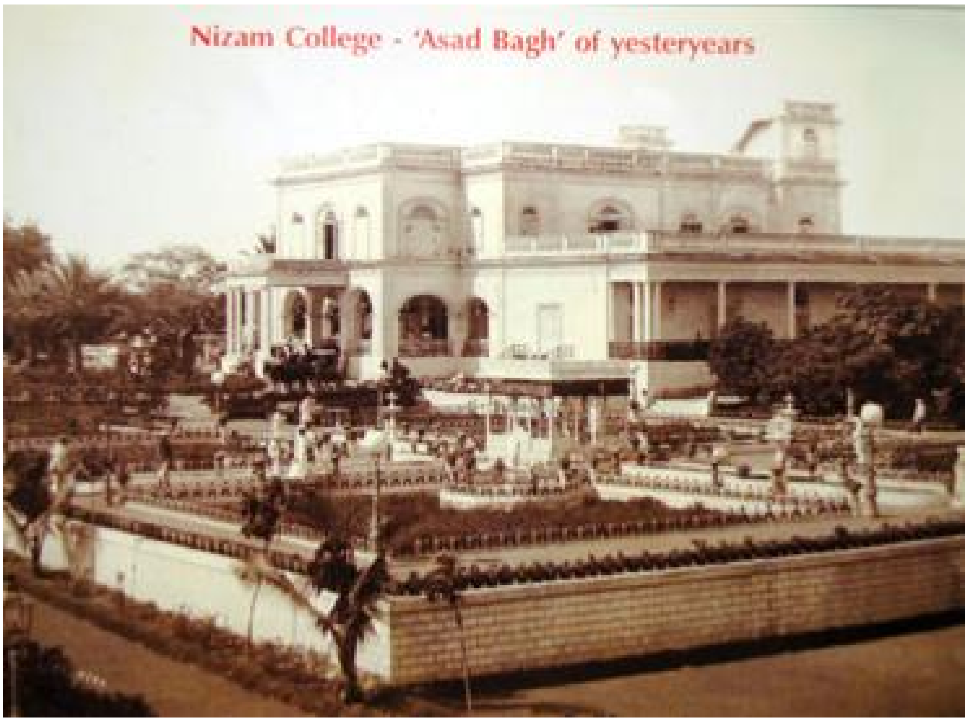
EAPCET
CHEMISTRY



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CHEMISTRY

By

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EAPCET

(Engineering, Agriculture & Medicine Common Entrance Test)

CHEMISTRY

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PREFACE

Taking competitive examinations has became the order of the day for any educated young man who is desirous of seeking any coveted job, a seat in any prestigious college. The approach required for such competitive examination is different from that of taking an academic examination.

It was observed that most of the minority candidates do not fare well at these competitive examinations not because they lack in talents but because they can neither afford to join the private coaching centres nor could purchase the required study material.

In order to improve the participation and performance of the candidates belonging to minorities in such competitive examinations, the Minorities Welfare Department, State Government sponsored a project to Osmania University. The University in turn established Centre for Educational Development of Minorities (CEDM) in 1994 in Nizam College. Since then, the Centre has been offering free coaching for the benefit of candidates belonging to minority communities appearing for various job seeking and admission seeking competitive examinations at Hyderabad and other minority concentrated districts of the state. In respect of job-seeking examinations, the Centre is providing free coaching and study material for TS TRT, TS TET etc. and for admission oriented examinations such as NEET, EAPCET, ICET, ECET, EdCET, DEECET and POLYCET etc. In addition to these coaching programmes, the Centre is also providing free coaching and study material to X class Urdu medium minority students in minority concentrated districts of the state to strengthen their educational foundation and to improve their performance in SSC Public Examination.

We wish to place on record the pains the compilers have taken to summarize and arrange the important questions. The Centre gratefully acknowledges their services.

If these study materials are of any help to the candidates, we feel immensely rewarded for the humble efforts we have put in.

Hyderabad
April 2024

Prof. S. A. Shukoor,
DIRECTOR

CONTENT

First Year

Secon Year

I YEAR

1. ATOMIC STRUCTURE

Synopsis:

1. The smallest indivisible particle of the matter that participates in chemical reaction and which may or may not exist freely is defined as atom.
2. A pure substance that contains the atoms of same kind is called an element. The term element was proposed by Robert Boyle.
3. The present system of symbols to the element was introduced by Berzelius.
4. The smallest particle of matter which has free existence is called molecule.
5. Atom consists of fundamental particles like electron, proton & neutron.

Particle	Mass	Charge(in Col)	e.s.u	Discoverer
Electron	9.109×10^{-31}kg 0.000548 amu.	-1.602×10^{-19}	$+4.802 \times 10^{-10}$	Thomson
Proton	1.670×10^{-27} kg 1.00728 a.m.u	$+1.602 \times 10^{-19}$	-4.802×10^{-10}	Rutherford
Neutron	1.674×10^{-27} kg 1.008665 a.m.u	–	–	James Chadwick

6. The heaviest fundamental particle is neutron, & the lightest particle is electron.
7. Atomic number (*Z*) is the number of protons. It is also equal to number of electrons.
8. Mass number (A):- No. of neutrons + No. of protons. Number of neutrons *A*-*Z*.
9. X rays are produced when cathode rays impinge on a metal target. Moseley found that the frequency of x- rays is characteristic of element.
10. Moseley law is $\sqrt{\vartheta} = a(Z - b)$, where ϑ = frequency of X-rays. Z = atomic number; *a* and *b* are constants.
11. J.J. Thomson model: - The atom is a spherical structure. It resembles the water melon red spongy mass represents protons the negative charge are embedded in it like the seeds.
12. Light is a form Electro magnetic radiation.
13. Wave length (λ):- It is the distance between two successive points in a wave (two crests or two troughs) is called as wave length it is denoted byλ. $\lambda = \frac{c}{\nu}$ Units – A^0, nm, cm & m
14. 1nm = 10^{-9}m =10^{-7}cm; 1A^0 = 10^{-8}cm = 10^{-10}m.
15. Frequency (ν):- The number of waves passing through a given point in one second is called the frequency.
16. Units: cycles per sec or hertz $\nu = \frac{c}{\lambda}$, $\nu \propto E$
17. Wave number: The number of wave spread in unit area. It is denoted by ν. $\nu = \frac{1}{\lambda}$. Units of ν are cm^{-1} or m^{-1}.

18. Amplitude: It is the maximum displacement of the particles of the wave from their mean position. It is a measure of intensity of color.
19. Velocity (C):- The distance covered by the wave in a given time. $v = \frac{C}{\lambda}$. Units - cm / sec.
20. Electromagnetic spectrum: It is the arrangement of different electromagnetic radiations in the decreasing order of their energies.
21. Cosmic rays, γ-rays, X-rays, U.V rays, visible region, I.R. rays, Micro waves, Radio waves.
22. When solids or liquids are heated to incandescence they emit light and give continuous spectrum.
23. When a gas is heated to a high temp it shows line spectra by emission of light.
24. Line spectra are characteristic of atom (atomic spectra).Band a spectrum is given by molecule.
25. Absorption spectrum is given by gas molecules due to absorption of light radiation's it shows dark lines on bright background.
26. The solar spectra contain dark lines due to absorption spectrum of hydrogen and helium. These are called as Fraunhoffer lines.
27. Emission spectra as given by atom due to emission of radiations it as bright lines on dark background.
28. Sodium shows two lines in emission spectra at 5890-5896° A. these are yellow lines. Every element has its characteristic spectra.
29. Hydrogen gives red lines in spectra. Its spectra are the simplest of all.
30. Hydrogen spectra show 5 series of lines.

Series:	Lyman	Balmer	Paschen	Brackett	Pfund
Region:	UV	Visible	N.I.R	I.R	F.I.R
n_1 :	1	2	3	4	5
n_2 :	2, 3...	3, 4...	4, 5...	5, 6...	6, 7...

31. Wave number of spectral line can be calculated by Rydbergs equation.
32. $\vartheta = \frac{1}{\lambda} = R\left(\frac{1}{n_1} - \frac{1}{n_2}\right)$, R = 109,678 cm^{-1} ($n_1 < n_2$); n_1 = lower energy level; n_2 = higher energy level.
33. Balmer series consist of four lines H_α, H_β, H_γ, H_δ. For H_α, $n_2 = 3$; H_β, $n_2 = 4$; H_γ, $n_2 = 5$; H_δ, $n_2 = 6$; ($n_1 = 2$).
34. Max Planck proposed that energy is emitted or absorbed by a body in form of discrete packets of energy called quanta. Energy of each quanta $E = h\nu$ Where ν is frequency h = Planck's constant 6.625 × 10^{-27} erg-sec or 6.625 × 10^{-34} J-sec.
35. According to Einstein's theory energy of these discrete packets called photons.
36. Photo electric effect: When light is incident upon metals like K and Rb they emit electrons. It is called as Photo electric effect. It was proposed by Einstein. $h\vartheta = W +$ K.E, where $h\vartheta$ = energy of photon W = work function and K.E = kinetic energy of emitted electron.

37. Rutherford gave solar or planetary model. The atom is mostly empty with a centrally located positively charged nucleus. This model shows electrons revolve around nucleus balanced by two forces, centripetal & centrifugal forces.
38. According to classical theory of physics the revolving electron should loose energy and fall in spiral path into nucleus and the spectra should be continuous. But atomic spectra is not continuous hence his model was not accepted.
39. Bohr postulate: Electrons revolve in certain, fixed discrete, circular paths. The energy of these orbits is fixed and known as stationary states. They are denoted with K, L, M, N....
40. In these shells the angular momentum of electron $mvr = \frac{nh}{2\pi}$.
41. If an electron moves from lower to higher energy level it absorbs energy.
42. If an electron jumps from higher to lower energy level emits energy. This energy absorbed or emitted is denoted by ΔE. $\Delta E = E_2 - E_1 = h\nu$. E_2 = Energy of higher level; E_1 = Energy of lower level.
43. If ΔE is large, light ray of higher frequency is emitted.
44. If ΔE is small, tight ray of smaller frequency is emitted.
45. Radius of orbit is proportional to square of principal quantum number. Radius of an orbit $0.529 \times n^2$ A^0or $0529 \times n^2$ 10^{-8} cm. First orbit has radius 0.529A°.
46. Radius of nucleus: 10^{-13} cm or 10^{-15} m (1 Fermi).
47. Velocity of electron in an orbit $V = \frac{2\pi Ze^2}{nh}$; $V \propto \frac{1}{n}$.
48. With Increase in 'radius of orbit the K.E of electron decreases but the potential energy increases and the total energy Increases.
49. Difference in the energies decreases as the distance from the nucleus Increases.
50. Energy of the electron in any orbit $E_n = -\frac{2\pi^2 Z^2 e^4 m}{n^2 h^2} = -\frac{313.6Z^2}{n^2}$ k.cal / mole

$= -\frac{13.6}{n^2}$ eV = $-\frac{2.18\times10^{-11}}{n^2}$ erg/atom.

51. When an electron is removed from an atom it is said to be at infinite, distance (n) from the nucleus. Its energy is zero.
52. When it jumps from infinite distance to a particular orbit its energy decreases from 0 to negative value.
53. When the electron jumps from any higher level to first level, Lyman series is observed. If electron jumps from 3, 4 levels to 2 level Balmer series is observed.

Higher level (n_2)	lower level (n_1)	Name of series
2, 3, 4, 5....	1	Lymen
3, 4, 5, 6....	2	Balmer
4, 5, 6, 7....	3	Paschen
5, 6, 7, 8....	4	Bracekett
6, 7, 8, 9...	5	Pfund

54. No. of spectral lines possible for the an electron, when it is excited to n orbit is $\frac{n(n-1)}{2}$.

55. No. of transitions = $\varepsilon^{\Delta n}$ or ε^{n2-n1}.If electron returns from 4th to 1st level $\varepsilon 4\text{-}1 = \varepsilon 3$.

56. Bohr's model explains the spectra of only one electron atoms like He^{+}, Li^{+2} etc.

57. It cannot explain Zeeman effect i.e., splitting of spectral lines in presence of magnetic field & stark effect i.e., splitting of spectral lines in presence of electric field.

58. Sommerfield's model: He proposed elliptical orbits in addition to circular orbits. There are sub-shell in main energy levels.

59. To denote the sub-level another quantum number called Azmithal quantum number is required. k can have values from 1 to n. k indicates number of sub-levels in energy level.

60. n / k = length of major axis / length of minor axis if then orbit is circular all the sub level in the same energy level have same (major axis) but differ in their shapes(minor axis).

61. Angular momentum of elliptical orbit $mvr = kh / 2\pi$. k is replaced by l which is known as angular momentum quantum no. or subsidiary quantum no. Values of l are from 0 to $(n - 1)$. l can have no. of values 0, 1, 2, 3... The sub shells are named as s, p, d, f.

62. Azimuthal quantum number indicates shape of orbital.
s - (sharp)–spherical
p - (principal)-dumb bell
d - (diffused)-double dumb bell
f - (fundamental)-complicated.

63. Debroglie Wave theory**:** If the two ends of electron wave meet to give a regular series of crests and troughs, around the nucleus in circular orbit, the electron wave is in phase. It behaves as a STANDING WAVE (or) non-energy radiating wave (or) stationary wave. The constructive interference of electron wave is observed. When an electron is in the phase, the circumference of Bohr's orbit ($2\pi r$) is equal to integral multiplies of wavelength (λ) of electron wave. $2\pi r = n\lambda$

$\therefore \lambda = \dfrac{2\pi r}{n}$, Here r = radius of orbit; n = integer 1, 2, 3, 4, -----

When the circumference of Bohr's orbit ($2\pi r$) is bigger (or) smaller than $n\lambda$, then the electron waves are said to be out of phase. Hence a destructive interference of waves occurs causing radiation of Energy. So the orbit cannot exist. Bohr's theory and de Broglie's theory are in perfect agreement with each other.

64. Heisenberg's Uncertainty Principle: "Simultaneous and exact determination of the position and momentum of sub-atomic particle like electron moving with high speed is impossible. If the uncertainty of position of small particle is Δx and uncertainty in momentum is Δp. then $\Delta x \cdot \Delta p \geq \dfrac{h}{n\pi}$ [here n = 1, 2, 3, 4-----]. For an electron revolving round the nucleus in an atom, the value of n is nearly 4. $\Delta x \cdot \Delta p \geq \dfrac{h}{4\pi}$. Electron must move with a velocity, 1835 times that of proton to have their wave lengths equal. For an electron if the uncertainty in the position is zero, then the uncertainty in the velocity will be infinite. Heisenberg uncertainty principle is against to Bohr's postulate of fixed.

65. Quantum Numbers: The quantum numbers designate position of electron in an atom.

Principal quantum number (n): It was proposed by Neil's Bohr. It represents radius, energy, and probability, of an electron in a given shell. As the value of n increases energy of an electron increases.

Energy in H atom in electron volts.

K – I orbit → – 13.6 ev
L – II orbit → – 3.4 ev
M – III orbit → – 1.5 ev

Total no of electrons present in a given shell is $2n^2$.

K $n = 1$ 2×1^2 2e-
L $n = 2$ 2×2^2 $8e^-$

Number of orbitals in a shell.

K $1^2 = 1$
L $2^2 = 4$
M $3^2 = 9$

Azimutal quantum number (l):

It was proposed by Arnold Sommerfeld. The value of Azimutal quantum number varies from 0 to (n–1). The total value is equal to principal quantum number value.

l represent the subshells

s -	sub shell	= 2 electrons	$l = 0$
p -	,,	= 6 electrons	$l = 1$
d -	,,	= 10 electrons	$l = 2$
f -	,,	= 14 electrons	$l = 3$

<u>Shell</u>		<u>Sub-shells</u>
n-1	$l = 0$	1s
n-2	$l = 0, 1$	2s ,2p
n-3	$l = 0, 1, 2$	3s, 3p, 3d
n-4	$l = 0, 1, 2, 3$	4s, 4p, 4d, 4f

Magnetic quantum number: It was proposed by Lande. Magnetic quantum number represents the orientation of orbital in space according to magnetic field or electric field. Magnetic quantum number explains Zeeman Effect and Stark Effect. The values of m varies from –l ....0......... + l the total value are equal to (2l+1). The orientation of different orbital represents the splitting of spectral lines into fine line due to Zeeman Effect and Stark Effect. Magnetic values depends upon l value

If $l = 0$ $m = (2 \times 0 + 1) = 1$ value is $m = 0$
If $l = 1$ m = $(2 \times 1 + 1) = 3$ values are $-1, 0 , +1$

s	$l = 0$	$m = 0$
p	$l = 1$	$m = -1, 0, +1$
d	$l = 2$	$m = -2, -1, 0, +1, +2, 3$
f	$l = 3$	$m = -3, -2, -1, 0,1, 2, 3$
g	$l = 4$	$m = -4, -3, -2, -1, 0, 1, 2, 3, 4$

Each value of magnetic quantum number represents one orbital containing capacity of two electron.

Spin quantum number: It was proposed by Ullenbeck and Goudsmidth. It describes the spin of an *e* on its own axis. The values of spin quantum number are +½ & –½. Clock wise direction +½, anti clockwise direction -½. When two electrons as paired spin, the spin value is 0. The electron with parallel spin are present in half filled state, they show greater stability due to minimum repulsion.

66. When the number of unpaired electron are present in an atom the magnetic moment is calculated by $\mu = \sqrt{n(n+2)}$ B. M. Where n = number of unpaired electrons. B.M. = Bohr's magnetron.

67. Shapes of orbitals: Orbital is the 90-95% probability of finding an electron in a three dimensional space around the nucleus. Ψ^2 = probability function. Ψ = Wave function. Eingen function of an electron is called an orbital. Orbit is 2 dimensional and orbital is 3 dimensional.

68. Schrödinger gave description of an electron in three dimensional spaces in the form of equation called Schrödinger wave equation.

$$\frac{\partial^2\Psi}{\partial x^2}+\frac{\partial^2\Psi}{\partial y}+\frac{\partial^2\Psi}{\partial^2}+\frac{8\pi^2 m}{h^2(E-v)\Psi}=0$$

69. Types of Orbitals:

s – Orbital:

1) It is spherically symmetrical. Non-directional orbital
2) 1*s* orbital is smaller than 2*s* orbital in size.
3) The probability distance of finding an *e*- in 1*s*-orbital is 0.529 A^0
4) Number of nodal surface for *s* - orbital is (*n*-1) or (*n*-l-1). Eg: for 1*s* orbital is 1-1=0; for 2*s* orbital is 2-1=1
5) For s-orbital l = 0, m = 0. The m values for s-orbital are 1.Value is zero and it can have one orbital.

p- Orbital:

1) Each *p*-orbital has 2 lobes. Based upon the orientation of lobes, the orbitals are named.
2) Each *p* –orbital has 1 nodal plane
   Px = *yz* nodal plane
   Py =*x*, *y* ,, ,,
   Py =*x*, *y* ,, ,,
3) Electron spends equal time in both the lobes. *Px*, *Py* and *Pz* orbitals are known as degenerate orbitals. *l* value =1, m = –1, 0, +1

d-Orbital: -

1) There are 5 *d*-orbitals.
2) Each d – orbital has 4 lobes with two nodal planes.
3) First 3 orbitals *dxy*, *dyz*, *dxz* have the lobes in between the axis. For *l* = 2, *m* = –2, –1, 0, +1, +2
4) $dx^2–y^2$ orbital has lobes on the axis *x* and *y*. dz^2 orbital has 2 lobes on the *z*-axis and a ring at the centre known as torus or smoke ring or collar.
5) *d*-orbital split into higher energy and lower energy orbitals in presence of ligand' or magnetic field. Lower energy orbitals *dxy*, *dyz*, *dzx* and $dx^2–y^2$, dz^2 are higher energy orbital.

f- Orbital:

1) They are complex shape or four folded dumbbell. It as eight lobes. It as three nodal regions and it can accommodate 14 electrons.
2) $l = 3, m = -3, -2, -1, 0, 1, 2, 3$

g-Orbital: $l = 4$, $m = -4, -3, -2, -1, 0, 1, 2, 3, 4$. No. of orbitals in $g = 9$, maximum number of electrons = 18.

70. Electronic configuration: - The distribution and arrangement of electrons in the main shells, subshells and sub-subshell [orbitals] of a ground state atom is called "Electronic configuration". The electronic configuration of atoms of elements is denoted by nl^x notation. Here n = principal quantum number =1, 2, 3, 4, ---- l = the orbital [azimuthal] quantum number and x = maximum number of electrons accommodated.

71. The maximum limit of x can be known from Pauli's principle. In the Electron – Box method, each orbital is shown by a box $\boxed{}$ and spin of electrons are shown by (↑) and (↓). Ex: *H*: $1s^1$ $\boxed{\uparrow}$

72. Aufbau Principle: According to this principle electron fill into various orbitals in the order of increasing energy of orbitals. Energy of orbitals is calculated by Moller's orbital energy level diagram and (n+l) rule.

(n+l) RULE: Lower the (n+1) value lower will be the energy of orbital. Higher (n+l) value, higher will be the energy. If both the orbitals have same (n+l) value, the orbital with lower n value is considered first.

$1s = 1 + 0 = 1$; $2s = 2 + 0 = 2$; $2p = 2 + 1 = 3$; $3s = 3 + 0 = 3$;

$3p = 3 + 1 = 4$; $3d = 3 + 2 = 5$; $4s = 4 + 0 = 4$

Among $2p$ and $3s$, the orbital $2p$ has lower energy due to lower n value. Among $3d$ and $4s$, the orbital $4s$ shows lower (n + 1) value.

73. Hund's rule of maximum multiplicity: It explains the stability of electronic configuration in the degenerate orbitals, according to which the degenerate set of orbitals should be filled with single electron first then the pairing starts. In *s*-orbital 2nd, in *p*-orbital 4th, in *d*-orbital 6th, in *f* orbital 8th electron starts pairing. No. of unpaired electrons can be calculated by studying of Hund's rule in degenerate orbitals. Ex: Nitrogen $1s^2, 2s^2, 2p^3$

$$\overset{1s^2}{\boxed{\uparrow\downarrow}} \quad \overset{2s^2}{\boxed{\uparrow\downarrow}} \quad \overset{2p^3}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$$

According to Hund's rule nitrogen contain three unpaired electrons in the ground state.

74. Pauli's exclusion principle: According to this principle no two electron in an atom can have the all the four quantum numbers values same. If third electron enters into the orbital it shows four quantum values same. According to Pauli's exclusion principle $1s^3$ configuration is not possible. According to above rule any orbital can accommodate only two electrons with opposite spin.

Eg: He - $1s^2$

	n	*l*	*m*	*s*
I e^-	1	0	0	+ ½
II e^-	1	0	0	– ½

75. Anamolous configuration: -

Cr - [Ar] $3d^5$, $4s^1$ $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}\quad\boxed{\uparrow}$

Cu -[Ar] $3d^{10}$, $4s^1$ $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\quad\boxed{\uparrow}$

Chromium – half filled; Copper – completely filled

Cr – configuration is stable due to half filled state.

Cu- Configuration is stable due to completely filled state

76. Lowering Energy or exchange energy:

Cr- [Ar] $4s^1$, $3d^5$ $\boxed{\uparrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$

This configuration is stable due to lowering of energy and greater exchange energy.

77. Iso – Electronic species: Iso-electronic species have same no of electron & different number of protons and different number of neutrons.

Examples:

C^{-4}	N^{-3}	O^{-2}	F^{-1}	Ne	Na^+	Mg^{+2}	Al^{+3}
6 + 4,	7 + 3	8 + 2	9 + 1	10	11 – 1	12 – 2	13 –3

Same electronic configuration – $1s^2$, $2s^2$, $2p^6$

78. Isosters: Molecules having same number of atoms and same number of electrons are called isosters.

Ex: N_2 & CO; CO_2 & N_2O; HCl & F_2; MgS & CaO

Question Bank:

1. Charge of one mole of electrons is:
(1) 96.5 coulombs
(2) 9.650 coulombs
(3) 96 coulombs
(4) 96,500 coulombs

2. The energy in ergs associated with a light ray having wave length 1A^0 is:
(1) 1.957×10^{-19}
(2) 1.9575×10^{-8}
(3) 1.9575×10^{-27}
(4) 1.9575×10^{27}

3. The frequency of a light ray having wave number 10^5 cm is:
(1) 9×10^{10} cycles/sec
(2) 9×10^{10} cycles/sec
(3) 3×10^{15} cycles/sec
(4) 3×10^{20} cycles/sec

4. The wave length of a light ray having frequency 10^{16} sec^{-1} is:
(1) 3000A^0 (2) 9000A^0
(3) 600A^0 (4) 300A^0

5. As the value of 'n' increases, the energy of the orbit:
(1) Decreases
(2) Increases
(3) Remains constant
(4) None

6. The number of protons, electrons and neutrons in HD^+ is:
(1) 2, 1, 1 (2) 3, 2, 1
(3) 3, 2, 2 (4) 2, 2, 1

7. A neutral atom has 2K, 8L, 8M and 2N electrons. The number of P electrons present in it are:
(1) 20 (2) 16
(3) 10 (4) 12

8. The four quantum number values for the last electron in an atom are $n = 3$, $l = 1$, $m = -1$, $s = ½$. The element to which that atom belongs is:
(1) B (2) Ga
(3) Al (4) In

9. The atomic and mass numbers of Pb are 82 and 208 and the atomic and mass numbers of Bi are 83 ands 209. Then the n/p ratio is:
(1) More in Pb
(2) More in Bismuth
(3) Equal in Pb & Bi
(4) None

10. The number of elliptical orbit in 5th orbit are:
1) 5 (2) 4 (3) 3 (4) 1

11. When an electron jumps from higher orbit to lower orbit, then the spectrum contain ______ lines on _____ back ground:
(1) Dark, bright
(2) Bright, dark
(3) Bright, bright
(4) Dark, dark

12. The wave length of an electron moving with a velocity $6.625x10^4$ cm/sec is:
(1)11000 (2) 14790
(3) 14079 (4) 97014

13. The radius of an electron in n^{th} orbit of hydrogen atom is given by the equation:
(1) $\frac{n^2h^2me^2}{4m^2}$ (2) $\frac{n^2h}{4\pi^2me^2}$
(3) $\frac{n^2h^2}{4\pi^2me^2}$ (4) $\frac{n^2h^2}{4\pi me^2}$

14. Which of the following has maximum frequency?
(1) U.V rays (2) I.R rays
(3) X-rays (4) Cosmic rays

15. the degenerate orbitals possess:
(1) Same energy
(2) Same size
(3) Same orientation
(4) All of them

16. The element 'A' has mass number 19 and atomic number 9 then the symbol of its ion is:
(1) A^- (2) A^+ (3) A^{-2}(4) A^{+2}

17. The lobes of *dxy*, *dzy* and *dxz* orbitals are situated:
(1) Along the axis
(2) In between the axis
(3) Above the axis
(4) Below the axis

18. The wavelength of a spectral line having energy 10^{-10} ergs is:
(1) 198.75A° (2) 1987.5A°
(3) 10875A° (4) 19.875A°

19. The ratio between the wave lengths of two light rays having energies 5.0×10^{-12} and 2×10^{-11} ergs is:
(1) 2 : 1 (2) 4 : 1
(3) 3 : 1 (4) 3 : 2

20. The region of wave length of visible light is:
(1) 3000 to 7600A°
(2) 3600 to 7600A°
(3) 3800 to 7600A°
(4) 3800 to 7800A°

21. The wave number of the spectral line having frequency 3×10^{14} sec^{-1} is:
(1) 10^4 cm^{-1} (2) $10^5 cm^{-1}$
3) 10^6 cm^{-1} (4) 10^{10} cm^{-1}

22. The energy associated with a Red light having wave length 7000A° is:
(1) 3×10^{-10} ergs
(2) 2.84×10^{-12} ergs
(3) 3×10^{-12}ergs
(4) 2.84×10^{12}ergs

23. The radius of 1st orbit of Hydrogen atom is:
(1) 5.29A° (2) 0.0529A°
(3) 0.529A° (4) 0.592A°

24. The specific charge of an electron is equal to coulombs/gm:
(1) 1.76×10^8 (2) 1.76×10^6
(3) 1.76×10^3 (4) 1.76×10^{10}

25. The total number of electrons in one molecule of CO_2 is:
(1) 16 (2) 20 (3) 22 (4) 8

26. The wave length of an electro magnetic radiation with energy 2.6×10^{-12} ergs is:
(1) 760A° (2) 76cm
(3) 7600cm (4) $7600A^0$

27. The ratio of the energies of photons having wave lengths 7000A° and 3500A° is:
(1) 1 : 2 (2) 2 : 1
(3) 0.5 : 2 (4) 2 : 0.5

28. Cl^- ion Ca^{+2} ions have identical
(1) Configuration
(2) Reactivity
(3) Electrode potential
(4) Change

29. The change in energy, when an electron shifts from one shell to the other shell is given by:
(1) $E = h / \vartheta$ (2) $E = h\vartheta / 11$
(3) $E = h\vartheta$ (4) None

30. The wave length of the light is 4×10^{-10} cm. Then its wave number is:
(1) $2.5 \times 10^8 cm^{-1}$
(2) $25 \times 10^8 cm^{-1}$
(3) $.25 \times 10^8 cm^{-1}$
(4) $250 \times 10^8 cm^{-1}$

31. In hydrogen atom,
(1) the K.E of electron = $+e^2/2r$
(2) the P.E of electron = $+e^2/2r$
(3) the total energy of electron $=-e^2/2r$
(4) the ratio of K.E and total energy of electrons is +1 : 1

32. Identify the correct statement
1) The number of lines is Lymen series of hydrogen is 1.
2) The number of lines in Lymen series of Li^{2+} ion are only 2.
3) The hydrogen atom is relatively stable when the electron is in the first quantum level
4) A sample of hydrogen atom are supplied with certain amount of energy, the electrons in the hydrogen atoms excite to the same level

The correct answer is
(1) 1 and 3
(2) 3 and 4
(3) 3 only
(4) all are correct

33. Carefully study the following statements
1) the total l values for a given n value are (l-1)
2) the shape of the electron cloud is given by azimuthal quantum numbers
3) a duplet is formed by the electronic transition of 2→1
4) the orbital angular momentum of an electron is given by $h\sqrt{l(l+1)}$

The correct statements are
(1) 1, 2 and 3
(2) 2, 3 and 4
(3) 2 and 3
(4) all are correct

34. Regarding degenerate orbitals some statements are given
1) degenerate orbitals have same n value
2) degenerate orbitals have same n and l-values
3) degenerate orbitals have same n, l and m-values
4) degenerate orbitals have similar size and orientation in space

The wrong statements are
(1) 1 and 3 (2) 2 and 3
(3) 3 and 4 (4) none

35. During compton effect it was observed that the wave length of incident x rays to the wave length of scattered X-rays is:
(1) Greater (2) Equal
(3) Lesser (4) None

36. According to quantum theory light is given out and propagate as:
(1) Quanta and waves
(2) Particles and waves
(3) Waves
(4) None

37. The d electron in 'M' shell is characterised by:
(1) $n = 4, l = 2$ (2) $n = 3, l = 2$
(3) $n = 2, l = 2$ (4) $n = 3, l = 3$

38. The ion that is iso-electronic with CO is
(1) CN^- $^-$(2) O_2^+
(3) O^{-2} (4) N_2^+

39. Assertion (A): Shapes of the orbitals are represented by boundary surface diagram of contrast probability density.
Reason (R): Boundary surface diagram helps in interpreting and visualizing an atomic orbital.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

40. Assertion (A): Total number of electrons in a subshell is designated by azimuthal quantum number is 2(2l+1).
Reason (R): l can have value 0, 1, 2, … n−1 where n is a principal quantum number.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**

(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

41. Assertion (A): $\mathrm{Cl}^{\surd}$ and $\mathrm{K}^{\oplus}$ ions have same electronic configuration.
Reason (R): $\mathrm{K}^{\oplus}$ and $\mathrm{Cl}^{\surd}$ ions are iso-electronic species.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

42. Assertion (A): $E = mc^2$ according to Plank's quantum theory.
Reason (R): The radiant energy will be transmitted in the form of waves according to Plank's quantum theory.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

43. Assertion (A): Hund's rule is not applicable for the elements H, He, Li, Be.
Reason (R): Degenerate orbitals are not present for those elements in valence shells.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

44. Observe the electronic configuration of same elements in ground state.
i) Cr : $5s^1\ 3d^5$ ii) Mn : $4s^2\ 3d^5$
iii) Cu : $4s^1\ 3d^{10}$ iv) Zn $4s^2\ 3d^{10}$
(1) i and iii (2) ii and iv
(3) ii and iii (4) all of these

45. Identify the wrong statements from the following.
1) Aufbau means building up
2) If two orbitals have same $(n+1)$ value then the election enters into the orbital with lower n value
3) The energy of $3d$-orbital is less than that of $4s$.
4) The energy of an orbital can be determined by n value.

The correct answer is
(1) 1 and 3 (2) 2 and 3
(3) 1 and 4 (4) 3 and 4

46. If Aufbau rule and Hund's rule are not considered which of the following statements are correct.
1) Fe^{2+} will have configuration as [Ar] $3d^2\ 3d^2\ 3d^2$
2) Cu^{2+} is colorless ion
3) Magnetic moment of the Mn is $\sqrt{3}$ BM
4) $\mathrm{K}^{\oplus}$ is of d-block

The correct answer is
(1) 1 and 2 (2) 2 and 3
(3) 1, 3 and 4 (4) all are correct

47. Match the following.

List I ***(Ions)***	**List II** ***(Magnetic Moment)***
A) Ti^{3+}, Cu^{2+}	1) 1.73 BM
B) Ti^{2+}, Ni^{2+}	2) 2.84 BM
C) Cr^{2+}, Fe^{2+}	3) 4.8 BM
D) V^{2+}, Co^{2+}	4) 3.8 BM
	5) 2.18 BM

The correct matching is

	A	B	C	D
(1)	4	3	3	1
(2)	1	2	3	4
(3)	5	3	2	1
(4)	4	5	3	2

48. Match the following.

List I ***Principle / Theory***	**List II** ***Expression related***
A) Quantum Theory	1) $\lambda = \frac{h}{mv}$
B) Einstein's equation	2) $E = mc^2$
C) De-Broglie's equation	3) $2\pi r = n\lambda$
D) Constructive Interference	4) $E = \frac{hc}{\lambda}$
	5) $2\pi r = \frac{n}{\lambda}$

The correct match is

	A	B	C	D
(1)	5	1	3	4
(2)	4	2	1	3
(3)	4	2	3	1
(4)	5	4	2	3

49. Match the following.

List I ***Element***	**List II** ***Quantum numbers values for differentiating electron***
A) K	1) 2, 1, 1, + ½
B) Be	2) 2, 1, 1, – ½
C) N	3) 3, 1, 0, – ½
D) Cl	4) 4, 0, 0, + ½

The correct match is

	A	B	C	D
(1)	4	5	3	2
(2)	5	2	1	3
(3)	3	4	5	1
(4)	2	3	4	5

50. What is the wavelength of a particle of mass 6.62×10^{-29} g moving with a velocity of 10^3 m/sec ($h = 6.62 \times 10^{-34}$ Js)

(1) 6.62×10^{-4} (2) 6.62×10^{-3}

(3) 10^{-5} (4) 10^5

KEY

1) **4**	2) **2**	3) **3**	4) **4**	5) **2**	6) **1**	7) **4**	8) **3**	9) **1**	10) **2**
11) **2**	12) **1**	13) **3**	14) **4**	15) **1**	16) **1**	17) **2**	18) **1**	19) **2**	20) **3**
21) **1**	22) **2**	23) **3**	24) **1**	25) **3**	26) **4**	27) **1**	28) **1**	29) **3**	30) **2**
31) **2**	32) **2**	33) **2**	34) **3**	35) **3**	36) **1**	37) **2**	38) **1**	39) **1**	40) **2**
41) **1**	42) **4**	43) **1**	44) **1**	45) **4**	46) **3**	47) **2**	48) **2**	49) **2**	50) **3**

2. PERIODIC CLASSIFICATION

Synopsis:

1. The total numbers of element presently in a periodic table are 109 from Hydrogen (H) to Metnerium (Mt).
2. There are 88 elements which are naturally available.
3. The first attempt to classify the elements was done by Dobernier by law of Triads.
4. According to law of triads if the elements are arranged in increasing order of their atomic weights the atomic weight of middle element is the average of two.
   Ex: Cl, Br & I   Li , Na & K   Ca, Sr & Ba.
5. Newland proposed the law of octaves.
6. Lother Mayer classified the elements based upon the atomic volume curves.
7. Halogens are present on the ascending portion of the curve, alkaline earth metals on the descending portion & transitional elements occupy the troughs of the curve.
8. First periodic table was constructed by Mendeleef. The periodic law is the physical and chemical properties are periodic functions of their atomic weights.
9. The gaps in Mendeleef's table were for missing elements, called 'eka' elements. Eka-boron was now discovered as Scandium, Eka-sodium as Germanium and eka-aluminium as Gallium.
10. Atomic weight of an element is the product of its equivalent weight and valency. Based on this atomic weights of Be and in are corrected.
11. Anomalous pairs are: 1. Cobalt-nickel, 2. Argon-potassium and 3. Tellurium-iodine.
12. Most of the atomic numbers were calculated by Mosely. Atomic number provided better basis for periodic arrangement of elements.
13. Isotopes as no place in the Mendeleef's table.
14. Modern periodic law is the physical and chemical properties of elements are periodic functions of their electronic configurations.
15. Long form of the periodic table is called Bohr table. Here the elements are arranged in the increasing order of atomic numbers.
16. There are 18 vertical columns called groups and 7 horizontal rows called periods.
17. Each period starts with an 'alkali metal' and ends with a 'noble gas'.
18. First period is the shortest period. It has only 2 elements. Second and third periods are short periods with eight elements each. Fourth and fifth periods are normal periods with eighteen elements each. Sixth periods are longest period with 32 elements. Seventh period is incomplete and it does not contain P-block elements.
19. The period number indicates the valence shell and the group number in Roman letters denotes number of electrons in the outer most shell. However, elements of zero group have eight electrons in the outer most shell.
20. Based on the entry of the differentiating electron, elements are classified into 4 blocks s, p, d and f-blocks.
21. *s*-block elements have ns^{1-2} configuration. They include alkali and alkaline earth elements (group IA and IIA)
22. *p*-block elements have ns^2np^6 configuration. The elements present are IIIA and VIIA groups and zero group. Many gases, non-metals and metalloids are present in the p-block.

23. The general configuration of *d*-block elements is $ns^{1-2}(n-1)d^{1-10}$. There are four series of *d*-block elements.
    3*d* series from Se (Z=21) to Zn (Z=30)
    4*d* series from Y (Z=39) to Cd (Z=48)
    5*d* series from La (Z=57) to Hg (Z=80)
    These series are present respectively in 4th, 5th and 6th periods with 10 elements each 6*d*-series is incomplete.
24. The general configuration of *f*-block elements is $ns^{1-2}(n-1)d^{0-1}(n-2)f^{1-14}$. There two series of *f*-block elements. 4*f*-series are called lanthanides, from Ca (Z=58) to Lu (Z=71). 5f –series are called Actinides, from Th(Z=90) to Lr(Z=103) Lanthanides are called rare-earths while actinides are mostly synthetic and radioactive.
25. Based on the electronic configuration and the properties, elements are classified into four types. They are typel: noble gases, type II: Representative elements, type III: transition elements, and type IV: inner transition elements.
26. Noble gases belong to zero group. They have the outer s and p orbitals completely filled. They are chemically inactive and are attributed to octet configuration. They exist as monoatomic gases.
27. Representative elements have outer shells partially filled. All *s* and *p* block elements except 'zero' group are representative elements. They are chemically active. They get the nearest inert gas configuration by losing or gaining or sharing electrons.
28. Elements in which two outer shells are partially filled are called transitions elements. They are d-block elements and general configuration is $ns^{1-2}(n-1)d^{1-10}$
29. Transition elements are all metals. They are hard with high melting, boiling points and densities. They are good conductors. They are all solids except Hg.
30. Transition metals have characteristic properties
    (a) variable oxidation states
    (b) colours in solution states
    (c) paramagnetism
    (d) forming alloys and complexes
    (e) catalytic activity
31. Elements of IIB, Zn, Cd and Hg are not considered as transition elements as they do not contain unpaired electrons in $(n-1)d$ subshell.
32. Elements in which three outer shells are partially filled are called inner transition elements. They are f-block elements.
33. Common oxidation state of transition elements is +2 and of inner transition elements is +3.
34. All the four blocks and four types of elements are present in the sixth period.

The naming of the new elements had been traditionally the privilege of the discoverer (or discoverers) and the suggested name was ratified by the IUPAC. In recent years this has led to some controversy. The new elements with very high atomic numbers are so unstable that only minute quantities, some times only a few atoms of them are obtained. Their synthesis and characterisation, therefore, require highly sophisticated costly equipment and laboratory. Such work is carried out with competitive spirit only in some laboratories in the world. Scientists, before collecting the reliable data on the new element, at times get tempted to claim for its discovery. For example, both American and

Soviet scientists claimed credit for discovering element 104. The Americans named it Rutherfordium whereas Soviets named it Kurchatovium. To avoid such problems, the IUPAC has made recommendation that until a new element's discovery is proved, and its name is officially recognized, a systematic nomenclature be derived directly from the atomic number of the element using the numerical roots for 0 and number 1-9. These are shown in ***Table.***

The roots are put together in order of digits which make up the atomic number and "ium" is added at the end. The IUPAC names for elements with Z and above 100 are shown in ***Table.***

Table IUPAC notation for naming elements with Z > 100

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	P
6	hex	h
7	sept	s
8	oct	o
9	enn	e

Table IUPAC Nomenclature of heavy elements

Atomic number	Root name of the element	Symbol of the element	Official name of the element	Official symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobehum	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassnium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Unn	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium*	Rg*
112	Ununbiium	Uub		
113	Ununtriium	Uut		
114	Ununquadium	Uuq		
115	Ununpentium	Uup		

* Official name and symbol yet to be announced by IUPAC

35. Thus, the new element first gets a temporary name, with symbol consisting of three letters. Later permanent name and symbol are-^given by a vote of IUPAC representatives from each country. The permanent name might reflect the country (or state of the country) in which the element was discovered, or pay tribute to a notable scientist. As of now, elements with atomic numbers up to 112, 114 and 116 have been discovered. Elements with atomic numbers 113,115,117 and 118 are not yet known.
36. Repetition of properties after intervals of z values 2, 8, 18 and 32 is called periodicity.
37. The cause for periodicity of properties is similar valence shell configuration.
38. Atomic radius is the distance between the nucleus and outer most electron of an atom. It is measured by X-ray diffraction techniques.
39. Atomic radius is of the order of 10^{-10}m. It is measured in A° or pm. Based on bonding is 3 types.
40. Covalent radius is one–half of the inter-nuclear distance between atoms of a homo diatomic molecule. It is used for non-metals.
41. Metallic radius or crystal radius is used for metallic elements. It is one-half of the inter-nuclear distance between adjacent atoms of a metal.
42. Vanderwaal's radius is used for molecular substances in the solid state. It is one-half of the inter-nuclear distance between atoms facing each other in adjacent molecules. It is 30-50% higher than the covalent radius.
43. Atomic radius depends upon the size of bonded atoms, the extent of attraction between atoms, bond order and the oxidation state.
44. Atomic radius increases down the group because of the formation of new shells.
45. In a period from left to right, atomic radius decreases as the nuclear charge increases
46. In any period, in general, atomic radius is least for halogen and highest for noble gas element. Radius is highest for the element 'Cs' and least for the element 'H'.
47. In transition elements, as the atomic number increases, the atomic radius decreases slightly. The decrease is insignificant because of screening effect by inner *d* orbitals.
48. Due to dispersed shape of *f*-orbitals and poor shielding effect, the atomic radius of lanthanides steadily decreases. This is called lanthanide contraction.
49. Due to lanthanide contraction, the crystal structure and other properties of lanthanides are similar and are difficult of to separate Lanthanide contraction is also observed it 5*d* elements. The size of 4*d* and 5*d* element in a group are almost similar.
50. The radius of cation is less than the parent atom and that of the anion is greater.
51. Ionization potentials are determined by discharge tube experiments.
52. Ionization potential values are dependent upon atomic radius, nuclear charge, screening effect, penetration ability of orbitals and electronic configuration.
53. As the nuclear charge increases the nuclear attraction on the outer electron increases and ionization potential increases.

54. The minimization of the nuclear attraction by the electrons present in the inner shells is called screening or shielding effect. When the screening effect increases, ionization energy decreases.
55. Extent of penetration of orbitals of valence electrons towards the nucleus is in the order: *s>p>d>f.* Ionisation energies are also in the same order.
56. Elements having completely filled or half filled configurations are mere stable and need more energy for ionisation.
57. In a group ionisation energy decreases, as the size and screening effect increases.
58. In a period ionisation energy increases as the nuclear charge increases and the size decreases. Among transition elements or inner transition elements of a series, variation in the ionisation energies is smaller.
59. In any period alkali metal has least and noble gas has highest ionisation energies. The element with highest first ionisation energy is Helium and Cesium is least.
60. In any period alkali metal has least and noble gas has highest ionisation energies. The element with highest first ionisation energy is Helium and least is smaller.
61. Halogens have high electron affinities. Electron affinity of noble gases is zero and alkaline earth metals is negative.
62. Electron affinity is highest for chlorine. It is less for fluorine; due to small size and stronger inter electronic repulsions.
63. In a period electron affinity gradually increases. In a group generally it decreases.
64. Electronegativity is the tendency of the atom of an element to attract the shared electron pair towards its nucleus.
65. Pauling's scale of electronegativities is based on bond energies. Bond energy of AB molecule is taken as the average of bond energies of A_2 and B_2 molecule Experiment bond energy of AB is more than the calculated value.
66. $$\Delta = E_{AB} - \frac{1}{2}\left(E_{AA} + E_{BB}\right)$$

    Δ is the difference in the bond energies

    $X_A - X_B = 0.208\sqrt{\Delta}$, where Δ is in Kcal mol^{-1}

    $X_A - X_B = 0.1017\sqrt{\Delta}$, where Δ is in KJ mol^{-1}
67. Taking the Electronegativity of hydrogen as 2.1.the values of other elements are determined.
68. Electronegetivity of fluorine is highest. Its value is 4. Values of O, N, and Cl are respectively 3, 5, 3 and 3. For noble gases the value is zero.
69. Mulliken's Electronegativity is the average of ionisation energy and electron affinity. Mulliten's scale is applicable for monovalent elements.
70. If the differences in electronegativity of bonded atoms are 1.7 or more, the compound formed is ionic. Electronegativity values are useful in writing chemical formulae and in calculation of oxidation states.
71. Number of hydrogen or chlorine atoms or twice the number of oxygen atoms with which one atom of the element combines is called valency.
72. Elements in a group show generally same valency. Across a period the valence increases unit by unit. Transition metals show variable valency. Maximum valency of an element is its group number.

73. Minimum valency is zero. Highest valency is exhibited by Os, Ru or Xe. The value is 8. In the second period highest valency of 4 is exhibited by carbon and in the third period highest valency of 7 is exhibited by chlorine.

74. Oxidation state is the possible charge which the atom of the element appears to have acquired in a given species.

75. Oxidation state could be positive or negative or zero. It may be fractional. It may be explained with the help of configuration and electronegativity.

76. Elements of s-block and fluorine exhibit fixed oxidation states. Osmium exhibits highest oxidation state of +8 and carbon lowest oxidation state of –4.

77. The tendency of an element to lose an electron is called electropositively or metallic nature. Alkali elements are highly electropositive.

78. Electropositive elements react with water and acids to liberate hydrogen Ions of strong electropositive metals do not undergo hydrolysis.

79. Electropositively increases down the group as the size increases and across a period decreases as the size decreases

80. Element with low electronegativity is a metal and with high electronegativity is a non-metal nature decreases down the group. Metallic nature decreases and non-metallic nature increases across a period. Best metal available is ‘Cs’ and non-metal is ‘F’.

81. Metallic oxides are generally basic and form alkaline solutions in water. Non-metallic oxides are generally acidic and form acidic solutions in water. Oxides of metalloids are normally amphoteric.

82. An element of a group of the second period has similar properties with the second element of the next group of the third period. This is called diagonal relationship.

83. The three diagonally related pairs of the elements are Li and Mg, Be and Al, B and Si. The reasons for the diagonal relationships are similarly in the electronegativity and similar polarizing power.

$$\text{Polarizing Power} = \frac{\text{ionic charge}}{(\text{ionic radius})^2}$$

84. Elements which exhibit diagonal relationship have similar properties. Oxides of Be and Al are amphoteric. Carbides of Be and Al produce CH_4 on hydrolysis.

Question bank:

1. The elements with similar chemical properties
   (1) contain same number of electrons in the outer orbit
   (2) are present in same period
   (3) possess same number of neutrons
   (4) possess same atomic weights

2. Which of the following element resembles the 10th element in the periodic table
   (1) First element (2) 18th element
   (3) 20th element (4) 9th element

3. The following group of elements do not possess equal number of valence electrons
   (1) 0 group (2) IA group
   (3) VIIA group (4) IIA group

4. Bridge elements are the elements of ___________ period
   (1) 1st (2) 2nd
   (3) 3rd (4) 6th

5. Typical elements are the elements of ___________ period
(1) 1st (2) 2nd
(3) 3rd (4) 5th

6. Ce-58 is a member of
(1) *s*-block (2) *p*-block
(3) *d*-block (4) *f*-block

7. The elements with atomic numbers 2, 10, 18, 36, 54 and 86 are called
(1) Rare gases (2) Halogens
(3) Chalcogens (4) Pnicogens

8. The atomic number of an element with incompletely filled 3d-subshell is
(1) 30 (2) 29
(3) 21 (4) 19

9. 3d-subshell is completely filled in
(1) Ca (2) Cr
(3) Ni (4) Cu

10. Transition metals exhibit variable oxidation states. It is because of
(1) the smaller atomic radius
(2) the higher nuclear charge
(3) the energy difference between $(n\text{-}1)d$ and ns - subshells is very less
(4) high screening effect

11. The atomic numbers of Lanthanides are from
(1) 57 to 70 (2) 90 to 103
(3) 58 to 103 (4) 58 to 71

12. $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^3$ is the electronic configuration of
(1) Phosphorus atom
(2) Silicon atom
(3) excited Phosphorous atom
(4) excited Silicon atom

13. The atomic number of a larger atom is
(1) Z=17 (2) Z=16
(3) Z=18 (4) Z=19

14. The following element is a Lanthanide
(1) $Z = 57$ (2) $Z = 70$
(3) $Z = 73$ (4) $Z = 72$

15. In which block of the periodic table, a liquid non metal is present
(1) s-block (2) p-block
(3) d-block (4) f-block

16. Which is a Typical element
(1) Br (2) Fe
(3) Ce (4) Al

17. The starting element in the incomplete period of Long form of periodic table
(1) Cs (2) Th
(3) Fr (4) Rn

18. The atomic radius depends on
(1) the number of bonds
(2) the nature of the bonding
(3) the oxidation state
(4) All of these

19. The atomic radius cannot be measured exactly. It is because
(1) the boundary of nucleus is not clear
(2) the nucleus is very smaller
(3) the atom is very larger
(4) the electron cloud is well distributed

20. With an increase in the atomic number, the atomic radius in a group_______ and in a period______
(1) increases and increases
(2) decreases and decreases

(3) increases and decreases
(4) decreases and increases

21. The ions O^{2-}, F^-, Na^+, Mg^{2+} and Al^{3+} are isoelectronic. Their ionic radii show
(1) a significant increase from O^{2-} to Al^{3+}
(2) a significant decrease from O^{2-} to Al^{3+}
(3) an increase from O^{2-} to F^- and then decrease from Na^+ to Al^{3+}
(4) a decrease from O^{2-} to F^- and then increase from Na^+ to Al^{3+}

22. The number of elements in 5th period are
(1) 8 (2) 18 (3) 32 (4) 24

23. Which of the following alkali metals is smallest in size
(1) Rb (2) K
(3) Na (4) Li

24. The radius of La^{3+} [Z=57] is 1.06A°. Which of the following given values will be closest to the radius of Lu^{3+} [Z=71]
(1) 1.60A° (2) 1.40A°
(3) 1.06A° (4) 0.85A°

25. Among the following iso-electronic species, the ion with larger size is
(1) Sc^{3+} (2) Ti^{4+}
(3) P^{3-} (4) S^{2-}

26. The energy required for the removal of outermost electron from an isolated unipositive gaseous ion is called
(1) First Ionisation potential
(2) Second Ionisation potential
(3) Third Ionisation potential
(4) Fourth Ionisation potential

27. Which is less for an element
(1) I_1 (2) I_2
(3) I_3 (4) I_4

28. The I_1, I_2, I_3 and I_4 values of an element "M" are 120kJ/mole, 600kJ/mole, 1000kJ/mole and 8000kJ/mole. Then the formula of its sulphate is
(1) $M_3(SO_4)_2$ (2) MSO_4
(3) $M_2(SO_4)_3$ (4) M_2SO_4

29. The following process requires more energy
1) $M^+ \rightarrow M^{2+}+e$
2) $M^{2+} \rightarrow M^{3+}+e$
3) $M \rightarrow M^{+}+e$
4) $M^{3+} \rightarrow M^{4+}+e$

30. The I_1 of Potassium is 4.339 eV/atom. The I_1 of Sodium is
(1) 4.339 eV/atom
(2) 2.21 eV/atom
(3) 5.138 eV/atom
(4) 1.002 eV/atom

31. The I_1 values of Li, Be and C are 5.4 eV/atom, 9.32 eV/atom and 11.26 eV/atom. The I_1 value of Boron is
(1) 13.6 eV/atom
(2) 8.29 eV/atom
(3) 14.5 eV/atom
(4) 21.5eV/atom

32. As the alkaline earth metals (except Be) tend to lose their valence electrons readily they act as
(1) weak oxidising agent
(2) weak reducing agent
(3) strong oxidising agent
(4) strong reducing agent

33. The valence configuration of a *p*-block elements is ns^2np^6. Here *n* is

I) 2	II) 3
III) 4	IV) 0
V) 1	VI)7

The impossible values are

(1) I, II and III (2) IV and V
(3) IV, V & VI (4) II and V

34. Regarding inner transition elements same statements are given below.

I) Their common oxidation state is +3.
II) Transuanic elements are synthetic elements.
III) The f-electrons do not involve in the bond formation.
IV) They exhibit different physical and chemical properties due to similar configuration of n and (n-1) shells of these elements.

The wrong statements are

(1) I, II and III
(2) I, II, III and IV
(3) I and III only
(4) None

35. Match the following.

Element	***Valence configuration***
A) Li	1) $5s^1$
B) K	2) $4s^1$
C) Rb	3) $1s^1$
D) H	4) $2s^1$
	5) $3s^1$

The correct match is

	A	**B**	**C**	**D**
(1)	5	2	1	3
(2)	4	2	5	3
(3)	4	2	1	3
(4)	1	2	3	4

36. Match the following.

Starting element of the period	***Last element of the period***
A) H	1) He
B) Na	2) Rn
C) K	3) Ar
D) Cs	4) Kr
	5) Xe

Which of the following is a wrong match

(1) A-1 (2) B-3
(3) C-5 (4) D-2

37. Match the following.

Element	***Valence configuration***
A) *s*-block	1) $ns^{1\rightarrow2}$
B) *p*-block	2) $ns^{1\rightarrow2}d^{1\rightarrow10}$
C) *d*-block	3) $ns^2np^{1\rightarrow16}$
D) *f*-block	4) $(n\text{-}2)f^{1\rightarrow14}(n\text{-}1)d^{0\rightarrow1}ns^2$
	5) $ns^{1\rightarrow2}(n\text{-}1)d^{1\rightarrow10}$

The correct match is

	A	**B**	**C**	**D**
(1)	1	2	3	4
(2)	4	3	2	5
(3)	1	3	5	4
(4)	5	4	3	2

38. Match the following.

Shell completed	***Element***
A) K-shell	1) H
B) L-shell	2) He
C) M-shell	3) Ar
D) N-shell	4) Ne
	5) Kr

The correct match is

	A	**B**	**C**	**D**
(1)	4	5	2	1
(2)	5	2	1	4
(3)	2	4	1	5
(4)	2	4	3	5

39. Assertion: Fe^{3+} is more stable than Fe^{2+} ion.
Reason: Fe^{2+} ion has half filled *d*-orbitals.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

40. Assertion: The physical and chemical properties of Inner transition elements are similar.
Reason: Inner transition elements show similar configuration in nth and $(n\text{-}1)^{th}$ shells.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

41. Assertion: Anion has larger size than that of atom.
Reason: Anion has smaller nuclear charge.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

42. Assertion: Electron affinity of oxygen is less than nitrogen but more than that of fluorine
Reason: N, O, F belong to same period of the periodic table
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

43. Assertion: Ground state configuration of Cr is $3d^5\ 4s^1$.
Reason: A set of half filled orbitals containing one electron each with their spin parallel provides extra stability to the system.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

44. Regarding Mendeleev's periodic table four statements are given.
I) The elements are arranged in the increasing order of atomic numbers.
II) The atomic weights of Be and I_n were corrected by Mendeleev.
III) Te, -I, Ar-K and Co-Ni are the anomalous pairs.
IV) The product of valency and equivalent weight of that elements is its atomic number.
The correct statements are
(1) I and IV (2) II and III
(3) I, II and III (4) I, II, III, IV

45. The atomic numbers of the elements present in the same group are
(1) 3, 11, 19, 37, 55 and 87
(2) 2, 8, 18, 20, 28
(3) 2,3,4,5,6 and 7
(4) 2,8, 16, 13 and 13

46. A transition metal ion M^{3+} has the configuration $1s^2 2s^2 2p^6 3s^2 3p^6$. Then that element “M” is
(1) Sc (2) Ti
(3) V (4) Cr

47. The number of unpaired electrons in Fe^{3+} ion will be
(1) 3 (2) 3
(3) 4 (4) 5

48. In an element “X” the differentiating electron enters into s-orbital and its configuration is $1s^2$. Then the next element is present in
(1) s-block (2) p-block
(3) d-block (4) f-block

49. Which of the following group contains two different block elements
(1) VIII (2) IIIB
(3) VIB (4) VIIB

50. An element with $Z = 27$ belongs to VIII group. The elements with $Z = 26$ and 28 belongs to ______ and _______ groups.
(1) VIIB and IIB
(2) VIB and VIIB
(3) IB and IIIB
(4) VIII and VIII

51. The electronic configuration of element Z=117 is
(1) $[Rn]7S^2 5f^{14} 6d^{10} 7p^5$
(2) $[Rn]7S^2 5f^{14} 7p^5$
(3) $[Rn]5f^{14} 7p^5$
(4) None of the above

52. The symbol of element is Z=120 is
(1) Uuo
(2) Uun
(3) Ubn
(4) Noe of these

KEY

1) **1**	2) **2**	3) **1**	4) **2**	5) **3**	6) **3**	7) **1**	8) **3**	9) **4**	10) **3**
11) **4**	12) **4**	13) **4**	14) **2**	15) **2**	16) **4**	17) **3**	18) **4**	19) **4**	20) **3**
21) **2**	22) **2**	23) **4**	24) **4**	25) **3**	26) **2**	27) **1**	28) **3**	29) **4**	30) **3**
31) **2**	32) **4**	33) **3**	34) **4**	35) **3**	36) **3**	37) **3**	38) **4**	39) **1**	40) **1**
41) **1**	42) **4**	43) **1**	44) **2**	45) **1**	46) **1**	47) **4**	48) **1**	49) **2**	50) **4**
51) **1**	52) **3**								

3. CHEMICAL BONDING

Synopsis:

1. The force of attraction that keeps the various atoms of a molecule together is called chemical bond.
2. Atoms try to get stability by lowering their energy and by attaining the electronic configuration of the nearest inert gas.
3. Electronic theory of valency:
    (i) It was proposed by Kossel and Lewis.
    (ii) Only the valence electrons take part is bond formation.
    (iii) Atoms or ions with eight electrons in their outer most shell are highly stable.This is called octet rule.
    (iv) Except He ($1s^2$) all other gases have ns^2np^6 (octet) configuration.
    (v) Bond formed by the transfer (loss or gain) of electrons is called electrovalent or ionic bond.
    (vi) Hydrogen and Lithium try to get the stable configuration ($1s^2$)
    (vii) The three major types of chemical bonds are
        a) Ionic b) Covalent c) Dative or Co-ordinate covalent
4. **Ionic or Electrovalent Bond:**
    (i) It was proposed by Kossel. It is formed by the transfer of electrons form highly electropositive element to a highly electronegative element.
    (ii) For ionic bond formation, the EN difference between the atoms should be 1.9.
    (iii) The electrostatic force of attraction existing between oppositely charged ions is called ionic bond.
    (iv) The bond formed between atoms of low IP and high BA is ionic bond.
    (v) Ionic bond is always formed between a metal and a non-metal.
    (vi) All alkali metal compounds are IONIC.
    (vii) Except Beryllium, alkaline –earth metals form ionic compounds.
    (viii) Cations of greater size and anions of small size favor the formation of ionic bond.
    (ix) When an atom is converted into a cation, size decreases and when concerted into anion, the size increases.
5. Polarization and Fajan's Rules: The distortion or deformation of the electron cloud of an anion by a cation is called polarization. The tendency of the anion to get polarised is called its polarisability.The power of a cation to distort the electron cloud of anion is called its polarisation power.
    (i) Fajan's rules describe the conditions for polarization.
    (ii) Smaller the cation, greater is its polarising power. ($Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$)
    (iii) Larger the size of anion, greater is its polarisability. ($I^- > Br^- > Cl^- > F^-$)
    (iv) Polarising power of a cation increases with its charge. ($Al^{+3} > Mg^{+2} > Na^+$)
    (v) Cation with Pseudo inert gas configuration has greater polarising power than cation with inert gas configuration. ($Ag^+ > Na^+; Zn^{+2} > Ca^{+2}$)
    (vi) Greater the charge on the anion, greater is its polarisability. ($N^{-3} > O^{-2} > F^-$)

(vii) Most Favourable conditions for electrovalency.
a) Large cation b) Small anion
c) Small charge on both the ions.
(viii) Lithium iodide is the most covalent among lithium halides.
(ix) Among chlorides of Na, Mg and Al, least ionic is $AlCl_3$.
(x) AgCl is less ionic than NaCl (due to the pseudo inert gas configuration of $Ag+$).
(xi) CaF_2 is insolvable in water due to its high lattice energy.
(xii) LiF is ionic but BeF_2 is less ionic due to greater polarising power of Be.
(xiii) Stability of certain salts like AgI, PbI_2, HgI_2, and SnI_4 can also be explained with the help of Fajan's rule

6. The charge on a molecule is zero. In case of Polyatomic ion, the net charge is possessed by the ion as a whole and not by a particular atom. It is, however, feasible to assign a formal charge on each atom.

Formal charge is a factor based on a pure covalent bond formed by the sharing of electron peirs equally by neighbouring atoms. Formal charge of an atom in a polyatomic molecule or ion is defined as the difference between number of valence electrons of the atom in free state and number of electrons assigned to the atom in the Lewis structure.

Formal charge is denoted by Q_f and is given by $Q_f = N_A - N_M = N_A - N_{LP} - 1/2\ N_{BP}$

N_A is the number of valence electrons in the free atom,

N_M is the number of valence electrons belonging to the atom in the molecule,

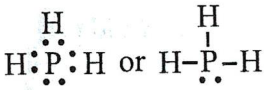
N_L is number of electrons in lone pairs and

N_B is number of electrons in bond pairs.

The counting of number of electrons is based on the assumption that the atom in the molecule owns one electron of each shared bonded pair and both the electrons of unshared lone pair.

Illustrative Example:

PH_3 molecule: The Lewis dot structure is



$$H:\overset{H}{\underset{\cdot\cdot}{\overset{\cdot\cdot}{P}}}:H \text{ or } H-\overset{H}{\underset{\cdot\cdot}{\overset{|}{P}}}-H$$

Formal charge of P:

$$Q_f = [N_A - N_M] = [N_A - N_{LP} - 1/2N_{BP}]$$
$$= \{5 - 2 - 1/2(6)\} = (5-5) = 0$$

7. **Covalent Bond:**

(i) It was suggested by Lewis.
(ii) It is formed by mutual sharing of electron pairs.
(iii) The number of electrons an atom shares with other is called covalency.
(iv) Covalency of hydrogen oxygen, nitrogen and carbon are 1, 2, 3 and 4 respectively.
(v) Number of covalent bonds formed between two atoms is equal to the number of electron pairs shared between them.
(vi) A single bond is formed by sharing one pair of electrons. H-H, Cl-Cl

(vii) Double bond by two pairs and triple bond by three pairs of electrons.
Ex : O=O, N≡N

(viii) covalent bond is of two types
(a) non-polar covalent bond.(Homo-atomic)
(b) polar covalent bond.(Hetro-atomic)

(ix) Non-polar covalent bond is formed by same atoms eg:H_2,N_2,O_2

(x) Polar covalent bond is formed by dissimilar atoms which differ significantly in their EN

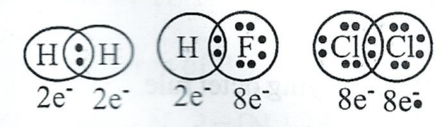
(xi) Polar covalent bond is formed by unequal sharing of electron pairs and is much stronger than non-polar covalent bond. Ex: $H^{+\delta}$-$F^{-\delta}$

8. Lewis dot model

Lewis dot model of bonding is a very simple model of explaining the structures of covalent molecules. Each valence electron of the atom involved in the bond formation is denoted by a 'dot'. Atoms share electrons in order to acquire ‘octet’ configuration. Hydrogen atom, however, acquires ‘duplet’ configuration. The ‘octet’ or duplet’ can be illustrated by drawing a circle around each atom.

Illustrative examples:

Hydrogen molecule is formed between two aydrogen atoms. By applying the same electron counting procedure to ‘F’ in HF molecule and to ' CF in CL molecule, we can write



By this scheme the ‘F’ atom in hydrogen fluoride and two ‘CF atoms in chlorine, all have eight electrons. ‘F’ has the neon configuration and ‘Cl’ has the argon configuration.

Lewis formulated the octet rule. Octet rule states that an atom must possess eight electrons in its outermost energy level for its stability. The octet rule also applies to molecules with more than one electron pair shared. The octet rule though useful, is not universal. It applies mainly to the second period elements.

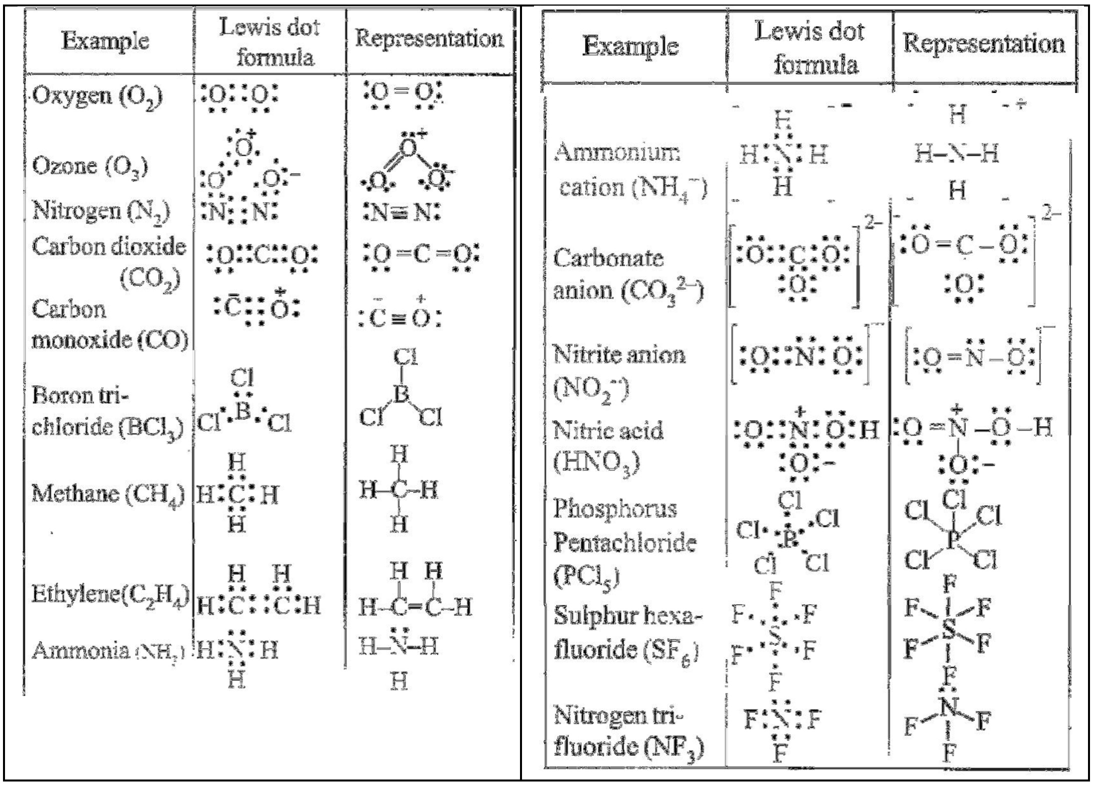
Exceptions to the octet rule

Incomplete octet: In some compounds, the number of electrons surrounding the central atom in a stable molecule is fewer than eight. Examples are $BeCl_2$, BF_3, $AlCl_3$, etc.

The expanded octet: There are more than eight valence electrons around the central atom in a number of compounds. This is termed as expanded octet. In addition to 3 s and 3p orbitals, elements in the third period also have 3d orbitals that can be used in bonding. Examples are SCl_4, PCl_5, SF_6, IF_7, H_2SO_4, etc.

Odd electron molecules : Some molecules contain an odd number of electrons. Examples are, NO, NO_2, ClO_2, etc. Octet configuration has even number of electrons. The octet rule clearly can. never be satisfied for molecules with odd number of electrons.

Cations : Some of the transition metal ions are known in which the octet is not complete. Examples are Fe^{2+}, Fe^{3+}, Co^{3+}, Ni^{2+}, Cu^{2+}, etc.



9. Valence Bond Theory (VBT):
   - (i) It was proposed by Heitler and London but improved by Pauling and Slater.
   - (ii) This theory is based on the principles of wave mechanics or quantum mechanics
   - (iii) Bond formation is favored when potential energy is minimum.
   - (iv) The bonding atoms are at optimum distance (Bond length)
   - (v) The bonding electron cloud is concentrated between the nuclei of the bonding atoms (located orbital)
   - (vi) Covalent bond is formed by the overlapping of atomic orbitals
   - (vii) Greater the extent of overlapping of atomic orbitals, greater is the strength of the bond.

Difference between a Sigma and a Pi bond

	σ-bond	**π-bond**
1) Type of overlapping	Formed by head on overlapping *s-s*, *s-p* or *p-p* overlapping	Formed by side wise overlapping Collateral) *p-p* (collateral)
2) symmetry in molecular orbital	Area of highest electron density along the inter nuclear axis	Two separate regions of high electron density , one each side of the axis
3) Region of Overlapping	Large	Small
4) Strength of bond	Strong	Weak
5) Formation of bond	Hybrid orbitals form	Formed only after a sigma bond is formed
6) Hybrid orbitals	Hybrid orbitals form a σ-bond	Hybrid orbitals cannot form a π-bond

(viii) VB theory can explain
  (a) directional nature of the covalent bond
  (b) bond angles (to some extent)
  (c) Shapes of molecules (simple molecules).
  (d) bond strength

(ix) VB theory or Lewis dot formula cannot explain the paramagnetic nature of oxygen molecule.

(x) The pair of electrons, belonging to the outermost shell of an atom, not involved in bond formation is called lone-pair of electrons.

10. **Hybridization**

(i) Hybridization was put forward by Pauling and it is a modification of VBT.

(ii) Spontaneous mixing of pure atomic orbitals, of the same atom of nearly equal energy to get the same number of identical new orbitals is called hybridization.

(iii) In hybridization
  (a) all the valence orbitals may or may not participate.
  (b) the hybrid orbitals differ in orientation but not in energy.
  (c) the hybrid orbitals always form strogent σ-bond than pure atomic orbitals.

(iv) *sp* hybridization:
  (a) two sp hybrid orbitals at 180° are formed
  (b) gives linear shape to the molecules
  (c) it is present in Beryllium halides, C_2H_2, CO_2, HCN etc.

(v) sp^2 hybridization:
  (a) Three sp^2 hybrid orbitals lying in a plane at 120° are formed
  (b) It gives trigonal planar structure to the molecule.
  (c) It is present in Boron halides, SO_2, SO_3, C_2H_4, $AlCl_3$ etc

(vi) sp^3 hybridization:
  (a) Four sp^3 hybrid orbitals at $109°28^1$are formed.
  (b) It gives tetrahedral shape to the molecule.
  (c) It is present in CH_4, SiH_4, CCl_4, $SiCl_4$, diamond silica, SiC etc

(vii) sp^3d hybrid orbitals:
  (a) trigonal bipyramidal in shape
  (b) angle between orbitals is 120° and 90°
  (c) present in PCl_5, PF_5 etc.,

(viii) sp^3d^2 hybrid orbitals:
  (a) octahedral in shape
  (b) angle is 90°
  (c) It is present in SF_6, SeF_6, TeF_6 etc.,

(ix) sp^3d^3 hybrid orbitals:
  (a) Pentagonal bipyramidal in shape
  (b) Angle between orbitals :72° and 90°
  (c) It is present in IF_7

11. Co-ordinate Covalent Bond (Dative Bond):
   - (i) It was proposed by Sidgwick
   - (ii) It is special case of covalent bond which the bond pair of electrons are furnished by one atom, but shared by both atoms.
   - (iii) The donar must contain at least one one pair of electrons.
   - (iv) The acceptor must posses at least one vacant orbital of suitable energy.
   - (v) Dative bond is present in NH_4^+, BF_2^-, H_3O^+, HNO_3, H_3PO_4, NH_3BF_3, all complex compounds etc.,

12. Compound with more than one type of bond

   1) NaOH → ionic ,covalent
   2) Na_2CO_3 → ionic, covalent
   3) KCN → ionic, covalent
   4) NH_4Cl → ionic, covalent, dative
   5) NH_3BF_3 → covalent, dative
   6) CO → covalent, dative
   7) $K_3[Fe(CN)_6]$ → ionic, covalent, dative
   8) $[Co(NH_3)_3Cl_3]$ → covalent, dative

13. Valence Shell Electron Pair Repulsion Theory (VSEPR)
   - (i) It was originally proposed by Sidgwick and Poweli later developed by Gillespie and Nylom.
   - (ii) It is applicable to simple molecules or ions of non-transitional elements.
   - (iii) Owing to repulsion among the electron pairs they are arranged in space around the central atom such that the repulsion are minimum.
   - (iv) Order of repulsion is lp – lp > lp – bp > bp > bp (lp =lone pair, bp = bond pair)
   - (v) Presence of lone pair of electrons on the central atom causes distortion to the regular shapes and deviation in bond angles.
   - (vi) If the central atom has 3 electron pairs of which one is lone pair ; the shape may be angular or bent or v-shape eg:$SnCl_2$
   - (vii) If the central atom has 4 electron pairs of which one is a lone pair; then the shape of the molecule is pyramidal eg:NH_3

14. **Dipole Moment** (Bond Polarity):
   - (i) Dipole moment (μ) is the product of charge (e) on one end of a molecule and distance between the opposite charges (d). $\mu = exd$
   - (ii) It is a vector quantity and is expressed in Debye units. 1D = 10^{-18} esu.cm
   - (iii) Molecules having opposite charges at their ends are described as dipoles and the bond between them is polar covalent.
   - (iv) As the difference in EN of the bonded atoms increases, polarity of the bond increases.
   - (v) Among the hydrogen halides, the most polar is HF and the least polar is HI in gaseous state

15. Dipole moment and its applications:
   - (i) Bond polarity can be predicated from dipole moment. Greater the value of dipole moment, greater is the polarity (ionic nature) of the bond.

     H-F 1.92D H-Cl 1.03D

(ii) Percentage ionic character of a covalent bond can be calculated

$$\% \text{ ionic character} = \frac{\mu \text{ observed}}{\mu \text{ calculated}} \times 100$$

(iii) The shape or symmetry of a molecule can be predicted. CO is linear because $\mu = 0$. SO_2 is bent because μ is not equal to 0.

(iv) Cis and trans isomers can be distinguished. μ of cis-isomer > μ of trans-isomer

(v) Higher solubility can be noticed O_3 is more soluble in water than O_2 due to its polarity.

(vi) Presence of a dative bond can be predicted

16. **Hydrogen Bond:**

(i) It was proposed by Latimer and Winmill.

(ii) The weak electrostatic force of attraction between the hydrogen atom of one molecule and highly electronegative atom (F, O, N) of another molecule or same molecule is called hydrogen bond.

(iii) Most favorable conditions for the hydrogen bond are (a) high electro negativity of the other atom and (b) small size of that atom.

(iv) Hydrogen bond is weaker than ionic , covalent, dative or metallic bond & stronger than dipole –dipole attraction or Vander wall's forces

(v) The strength of hydrogen bond is nearly 3 to 10 K.cal / mole

(vi) Hydrogen bond length is larger than normal covalent bond length

(vii) Order of strength of hydrogen bond is H---F > O--H---O > N---H---N

(viii) Strongest hydrogen bond is present in HF.

(ix) Although Cl and N have the same EN value, cannot form hydrogen bond due to its later size.

(x) Hydrogen bond is of two types (a) Intermolecular (between molecules) (b) Intermolecular (within the same molecule)

(xi) HF, H_2O, NH_3, lower alcohols (CH_3OH, C_2H_5OH), lower carboxylic acids (HCOOH, CH_3COOH), primary amines, secondary amines, p-nitro phenol, carbohydrates (glucose, fructose), DNA are examples of inter molecular hydrogen bonding

(xii) O-nitro phenol o-chlorophenol, salicyl aldehyde, salicylic acid are examples for

(xiii) Intra-molecular hydrogen bonding

(xiv) The following observations are due to inter-molecular H-bonding.

(a) H_2O boils at 100°C while H_2S at –66°C.

(b) C_2H_5OH has higher b.p. than dimethyl ether.

(c) CH_3COOH shows higher molecular weight due to association

(d) Formation of anions like HF.

17. **Metallic Bond:**

(i) The Bond formed between two electro positive elements are called metallic bond.

(ii) The force that binds the metal ion to the mobile electrons is called metallic bond.

Eg: bond present in metals such as Cu, Al, Fe etc

(iii) Strength of metallic bond increases with number of valence electrons decrease in size of the atom.

(iv) Electrical and thermal conductivity of metals is due to the free movement of delocalised electrons.

(v) Transition metals generally form stronger metallic bonds due to their small size and the participation of $(n-1)$ d electrons in bonding.

(vi) The following properties of metals can be explained by metallic bond.luster, malleability ductility, high conductivity, high melting and boiling points, opaqueness and toughness.

(vii) According to sea model of metallic bond, the cores of metal atoms are floating in a sea of electrons.

18. **Bond Length:** Average of the inter nuclear distance of the bonded atoms is bond length (expressed in A°or pm) $1pm=10^{-13}m$.Bond length decreases with increase in the number of bonds between the atoms. Pi bond formation shortens bond length. For a given atom, bond length increases with the size of the other atom bonded to it. HI > HBr > HCl > HF

19. **Bond angle:** The angle between the imaginary lines joining the nuclei of bonded atoms in molecules is called the bond angle. Bond angle depends on the state of hybridization and the repulsion between the electron pairs. Generally, greater the bond angle, stable is the molecule.

20. **Bond enthalpy:** Energy released when a covalent bond is formed from the isolated gaseous atoms is known as bond enthalpy. (expressed in Kcal/mol or KJ/mol)

    (i) The energy required to break the covalent bond is known as bond dissociation enthalpy.

    (ii) In case of diatomic molecules, bond enthalpy and bond dissociation energy are equal in magnitude but opposite in sign.

    (iii) In case of polyatomic molecules, the average of all the bond dissociation energies is taken as bond enthalpy.

Molecular orbital theory:- Hund and Mulliken postulated this theory.

- In this theory molecule is presumed to be a multic entered structure the electron reside in molecular orbitals instead of atomic orbitals.
- Molecular orbital is a region around nuclei of bonded atoms in a molecule where the probability of finding electrons is maximum.

Aspects:

- In a molecule, orbitals are associated with energy known as molecular orbitals.
- Molecular orbitals are formed by combination of atomic orbitals lineary.
- The number of molecular orbitals is equal to number of atomic orbitals combining.
- Shapes of molecular orbitals depend on the shapes of atomic orbitals.
- Molecular orbital accommodate a maximum of two electrons of opposite spin.
- The electrons of all the atoms in a molecule are revolving under the influence of all nuclei.
- The molecular orbitals are arranged in the increasing order of their energies; electrons are also filled in same order.

- Degenerate orbitals are of same energy the electrons get paired only when all are partially filled.
- If more atomic orbitals combine, one half have lower energy and other half have higher energy.
- The molecular orbitals with lower energy than Ao's are known as bonding, which are of higher energy anti bonding orbitals and are not involved are called as non-bonding orbitals.
- Order of energies, ant bonding>non-bonding>bonding orbitals.
- Bonding orbitals designated as σ, Π & anti bonding orbitals as $\overset{+}{\sigma} \& \overset{+}{\Pi}$
- When the sign of wave function of the atomic orbitals are same bonding molecular orbital formed.
- If it is different anti bonding molecular orbital formed.

FILLING OF ELECTRONS INTO MOLECULAR ORBITALS.

- Mo's are arranged in increased order of their energies. The lowest available Mo is filled first according to Afbau Principle.
- Mo can accommodate a maximum of two electrons with opposite spins (Pauli's exclusion followed)
- When more than one Mo of some energy is available, pairing taking place only after all the degenerate orbitals are partially filled.

The sequence energy levels of molecular orbitals is given by

$$\sigma 1S < \sigma 1S^{*} < \sigma 2S < \sigma 2S < \sigma 2px < \pi 2py < \pi 2pz \pi 2py \pi 2pz < \sigma 2px$$

after $\mathbf{14\bar{e}}$

$$\sigma 1S < \sigma 1S < \sigma 2S < \sigma 2S \pi 2py, \pi 2pz < \sigma 2px < \pi 2py, 2pz < \pi 2py, 2pz$$

Bond order = Bonding of electrons Anti bonding electrons
2

Eg: $N_2 \quad 7 \times 2 = 14\bar{e}$

$$\sigma 1S^2 \sigma 1S^2 \sigma 2S^2 \sigma 2S^2 \pi^2 2py, \pi^2 2pz \sigma 2px$$

Bond order $\frac{10-4}{2} = \frac{6}{2} = 3$

Eg: $O_2 \quad 8 \times 2 = 16\bar{e}$

$$\sigma 1S^2 \sigma 1S^2 \sigma 2S^2 \sigma 2S^2 \sigma 2px \sigma^2 2px, \pi^2 2py \pi^2 2px, \pi^1 2py \pi^1 2pz$$

Bond order = $\frac{10-6}{2} = \frac{4}{2} = 2$

Hence O1 is paramagnetic in nature.

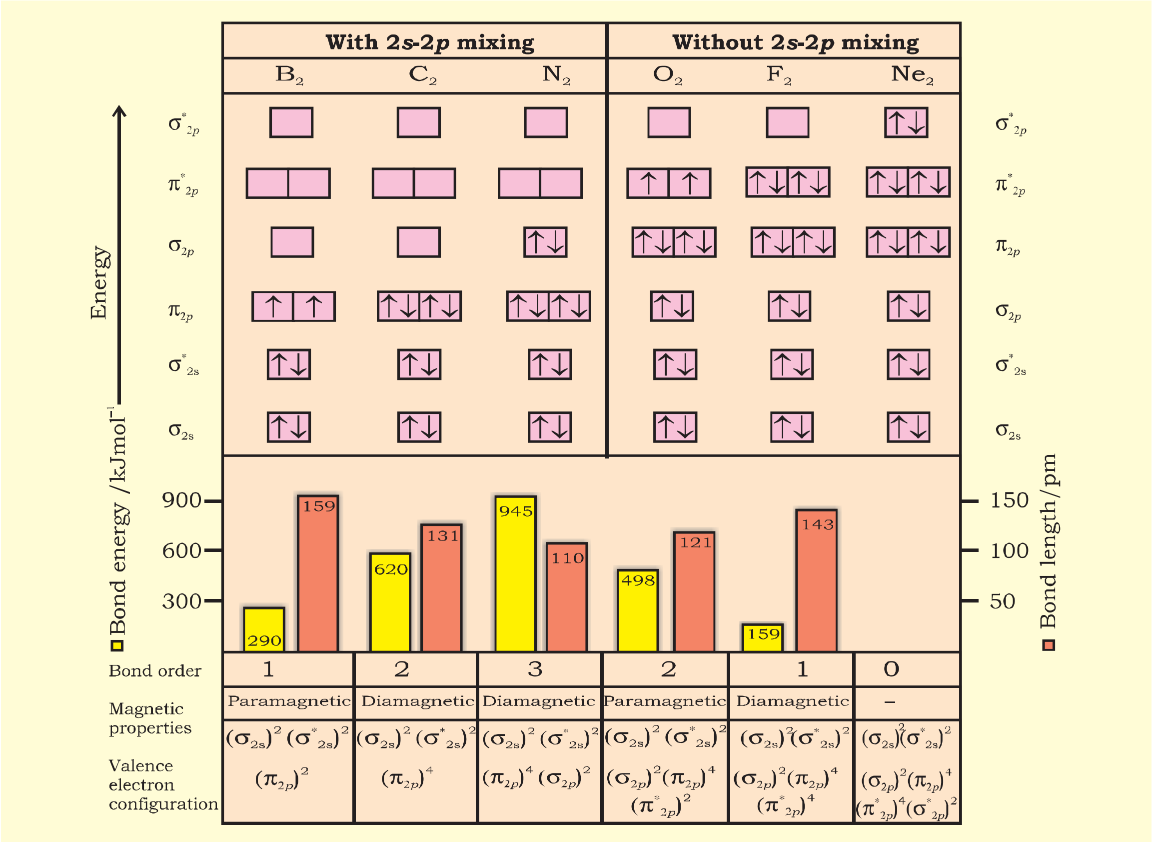


Fig. *MO occupancy and molecular properties for* B_2 *to* Ne_2.

Question bank:

1. The Configuration of an element X is $1s^1$.The Configuration of another element Y is $1s^22s^22p^5$. The bond formed between X and Y is
(1) Ionic (2) Covalent
(3) Metallic (4) Dative

2. Ionic bond is present in
(1) AlF_3 (2) $AlCl_3$
(3) $AlBr_3$ (4) HF

3. Ionic bond is stronger in
(1) KF (2) NaF
(3) CsCl (4) RbCl

4. The electronic configurations of the elements L,P,Q and R are given below
(1) L = $1s^22s^22p^4$
(2) Q = $1s^22s^22p^63s^23p^5$
(3) P = $1s^22s^22p^63s^1$
(4) R = $1s^22s^22p^63s^2$

5. What structural units occupy the lattice sites in the metallic crystals
(1) atoms
(2) electrons
(3) negative ions
(4) metal kernels

6. Which of the following is not a general property of covalent compounds
(1) they are soft
(2) chemical reactions are slow
(3) the smallest unit is molecule
(4) they are water soluble

7. Which one of the following is the most polar

(1) I-F (2) Br-F
(3) Cl-Br (4) Cl-F

8. Which of the following statements is true
(1) HF is less polar than Br
(2) Absolutely pure water does not contain any ions
(3) Chemical bond formation take place when forces of attraction overcome the forces of repulsion
(4) In covalency transference of electron takes place

9. Camphor is not soluble in
(1) Benzene (2) Ether
(3) Water (4) CCl_4

10. The contribution of ion present on the edge of the face of a unit cell is
(1) $\frac{1}{8}$ (2) $\frac{1}{2}$ (3) $\frac{1}{4}$(4) 1

11. Which of the following is least ionic
(1) AgCl (2) KCl
(3) $BaCl_2$ (4) $CaCl_2$

12. Which of the following is electron dificient
(1) BCl_3 (2) PCl_3
(3) PCl_5 (4) NH_3

13. Which of the following element exhibit same valency in all compounds
(1) S (2) Na
(3) Fe (4) Co

14. Which of the following is not paramagnetic
(1) NO (2) S^{-2}
(3) O^{2}_{-1} (4) O_2

15. Which is electron deficient compound
(1) C_2H_4 (2) B_2H_6
(3) C_2H_6 (4) $NaBH_4$

16. Camphor is not soluble in
(1) Benzene (2) Ether
(3) Water (4) CCl_4

17. Which of the following is true?
(1) Hybridisation of the central atom in NH_3 and CH_4 is sp^2
(2) $BeCl_2$ has "V" shape while SO_2 is linear
(3) SF_6 is octahedral and F-S-F bond angle is 90°
(4) CO_2 has dipole moment

18. Assertion: The bond energy of C≡C in acetylene is more than C=C in ethylene
Reason: As the multiplicity of bonds increases bond energy increases.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

19. Assertion: the electrovalency of magnesium is MgF_2 is two and that of Fluorine is one.
Reason: In the formation of ionic compounds, the number of electrons of atom of element loses (or) gains called electrovalency
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**

(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

20. Assertion: SF_6 obeys the octet rule
Reason: $BeCl_2$ is an example for electron deficient molecule
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

21. Assertion: In $NH^+{}_4$ ion, the four hydrogens become identical.
Reason: KCN contains covalent, dative and ionic bonds.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

22. Assertion: According to valence bond theory, the metallic properties vary from metal to metal.
Reason: These are varying according to their bond strengths and number of bonds in a lattice
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

23. Assertion: G.N. Lewis explains the formation of covalent bond.
Reason: Pure covalent bond is present in H_2 molecule.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

24. Average C-H bond energy is $416KJmol^{-1}$. Which of the following is correct?
(1) $CH_4(g)+416kJ \rightarrow C(g)+4H(g)$
(2) $CH_4(g) \rightarrow C(g)+4H(g)+416kJ$
(3) $CH_4(g) + 1664kJ \rightarrow C(g) + 4H(g)$
(4) $CH_4(g) \rightarrow C(g) + 4H(g) + 1664kJ$

25. Which of the following molecule contains one lone pair of electrons on the central atom?
(1) NH_3 (2) CH_4
(3) $CHCl_3$ (4) Cl_2

26. Ammonia has abnormally high boiling point because it has
(1) Alkaline nature
(2) Distorted shape
(3) sp^3 hybridisation
(4) Hydrogen bonding

27. The lusture of the metal is on account of the following
(1) High density of metals
(2) High polish of metals
(3) Chemical intertness of metals
(4) Reflection of light due to the presence of free electrons

28. The geometry of H_2S and its dipole moment are
(1) Angular and non-zero
(2) Angular and zero
(3) Linear and non-zero
(4) Linear and zero

29. Which of the following statements is true

(1) HF is polar than HBr
(2) Absolutely pure water does not contain any ions
(3) Chemical bond formation take place when forces of attraction overcome the forces of repulsion
(4) In covalency transference of electron takes place

30. Match the following.

List - I	*List-II*
a) C_2H_6	1) Ionic bond
b) Iron wire	2) Metallic bond
c) H_2O	3) Coordinate, covalent bonds
d) H_3O^+	4) Coordinate

The correct match is

	A	**B**	**C**	**D**
(1)	2	3	4	1
(2)	3	2	4	4
(3)	1	2	4	3
(4)	1	5	2	3

31. Match the following.

List – I	*List - II*
a) NH_4^+	1) SP^3 hybridisation, two lone pairs
b) H_3O^+	2) SP^2 hybridisation, one lone pair
c) XeO_3	3) SP^3 hybridisation, no lone pairs
d) SO_3	4) SP^2 hybridisation, no lone pairs

The correct match is

	A	**B**	**C**	**D**
(1)	1	2	4	5
(2)	2	2	3	5
(3)	3	4	4	5
(4)	4	4	3	5

32. Match the following.

List - I	*List - II*
A) Highly covalent material	1) Potassium chloride
B) Conductor of electricity	2) Ice
C) Good ionic solid	3) Diamond
D) Hydrogen bonded solid	4) Graphite
	5) Zeolite

The correct match is

	A	**B**	**C**	**D**
(1)	3	2	1	5
(2)	2	3	2	4
(3)	3	4	1	2
(4)	5	4	1	2

33. KCl is highly soluble in the following solvent

(1) C_6H_6 (2) CCl_4
(3) H_2O (4) C_2H_5-O-C_2H_5

34. Which of the following does not conduct electricity

(1) molten NaOH
(2) molten KOH
(3) solid NaCl
(4) aqueous NaCl

35. **A:** The rate of reaction among the covalent compounds is slow.

R: In ammonia boron trifloride, BF_3 is donor.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

36. **A:** Solid NaCl is a bad conductor of electricity.

R: In solid state the ions are immovable.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

37. **A:** A weak sigma bond is present in Cl_2 molecule.
R: The strength of *p-p* overlapping is more than *s-s* overlapping
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

38. **A:** VBT cannot explain the structure and bond angle in ammonia molecule.
R: Pure atomic orbitals are not involved in the formation of NH_3 molecule.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

39. **A:** The hybridization of carbon in $COCl_2$ is sp^2.
R: The shape of $COCl_2$ is linear.

40. **A:** NH_3 has pyramidal shape but $NH_4^{\oplus}$ has tetrahedral shape.
R: In NH_3 only 3 bond pairs and one lone pair is present but in $NH_4^{\oplus}$ 4 bond pairs are present.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

41. **A:** The bond angle in CH_4 is 109° 28' but in acetylene it is 180°.
R: With the increase of Y. S character in hybrid orbitals bond angle increases.

42. **A:** Though the bond length of F_2 is least the bond energy is also less.
R: Lone pairs of electron cause greater repulsions.
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

43. Study the following data.
I) Bond energy of HF is a
II) Bond energy of O=O is b
III) Bond energy of N≡N is c
IV) Bond energy of C–C is d
(1) $c > a > b > d$ (2) $c > b > a > d$
(3) $d > c > b > a$ (4) $a > b > c > d$

44. Which of the following is not correct?
I. $1A = 10^{-8}$ cm
II. $1nm = 10^{-9}$ m
III. $1Pm = 10^{-12}$
IV. $1nm = 10^{-14}$ cm
V. $1Pm = 10^{-8}$ cm
(1) I and V (2) II and III
(3) I, III & IV (4) IV and V

45. Pick out species having same kind of overlapping as an H_2 molecules from the following.

I) Cl_2	II) HCl
III) O_2	IV) Li_2
(1) I and II	(2) I and IV
(3) I, II, III & IV	(4) none

46. Increasing order of size of hybrid orbitals is

(1) sp, sp^2, sp^3	(2) sp^3, sp^2, sp
(3) sp^2, sp^3, sp	(4) sp^2, sp, sp^3

47. Dipole moment is shown by
(1) 1, 4 – dichloro benzene
(2) Cis 1, 2 – dichloro benznene
(3) Trans – 1, 2 – dichloro benzene
(4) Trans 2, 3 – dicholoro – 2-butane

48. Match the following.

List – I ***Compound***	List – II ***Nature***
A) Hydrogen Fluride	1) Monomer
B) Perchloric acid	2) Dimer
C) Acetic acid	3) Hexamer
D) Water	4) Polymer
	5) Trimer

The correct match is

	A	**B**	**C**	**D**
(1)	3	2	2	4
(2)	5	2	1	4
(3)	3	2	5	5
(4)	1	3	2	4

49. Match the following.

List – I	List – II
A) SO_3	1) sp^2
B) O_2	2) sp
C) NH_3	3) sp^3
D) PCl_5	4) sp^3d
	5) sp^3d^2

The wrong match from the List-I and List-II is

(1) A-1	(2) B-2
(3) C-4	(4) D-4

50. Match the following.

List – I ***Compound***	List – II ***Bonding***
A) NH_3 (liquid)	1) Inter molecular hydrogen bonding
B) Water	2) Intra molecular hydrogen bonding
C) CH_3COOH (liquid)	3) Dative bond
D) HF (liquid)	4) Ionic bond
	5) Covalent bond

Among them the wrong match is

(1) A-1	(2) B-5
(3) C-3	(4) D-1

51. Which of the following orders regarding the bond length is correct?
(1) $O_2^- > O_2 > O_2^+$
(2) $O_2^- < O_2 < O_2^+$
(3) $O_2^- > O_2 < O_2^+$
(4) $O_2^- < O_2 > O_2^+$

52. The molecular electronic configuration of B_2 is
(1) $KK(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p)_x^1(\pi 2p)_y^1$
(2) $KK(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p)_x^2$
(3) $KK(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p)^2$
(4) $KK(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p)^1(\pi 2p)^1$

53. Which of the following is paramagnetic with bond order 0.5
(1) O_2 (2) F_2 (3) N_2 (4) H_2^+

54. The formula for formal charge Q of an element in a molucle is
(1) $[N_A - N_{LP} - 1/2N_{BP}]$
(2) $[N_A - 1/2N_{BP}]$
(3) $[N_A - 1/2N_{LP}]$
(4) $[N_A - N_{LP}]$

55. The correct statement of Lewis Dot Symbol is

(1) On valance electrons are represented as dots on the symbol

(2) Total electrons are represented as dots on the symbol

(3) Only electrons as dots in the bonds are shown

(4) Only non bonding electrons are shown as dots

KEY

1) **2**	2) **1**	3) **2**	4) **3**	5) **4**	6) **4**	7) **1**	8) **3**	9) **2**	10) **3**
11) **1**	12) **1**	13) **2**	14) **2**	15) **2**	16) **2**	17) **3**	18) **1**	19) **1**	20) **4**
21) **3**	22) **1**	23) **2**	24) **3**	25) **1**	26) **4**	27) **4**	28) **3**	29) **3**	30) **3**
31) **3**	32) **3**	33) **3**	34) **3**	35) **3**	36) **1**	37) **4**	38) **1**	39) **3**	40) **1**
41) **1**	42) **1**	43) **1**	44) **4**	45) **4**	46) **1**	47) **2**	48) **1**	49) **3**	50) **3**
51) **1**	52) **1**	53) **4**	54) **1**	55) **1**					

4. STATES OF MATTER

Synopsis:

79. Matter exits in three different states i.e., solid, liquid and gas.
1. Gases have neither definite shape nor definite volume.
2. Gases can be liquefied at high pressure and low temperature.
3. Gases have high compressibility, low molecular weight and low density.
4. The measurable properties of gases are mass, volume, pressure and temperature.
5. The Boyle's law gives the relationship between the volume and pressure of gas at constant temperature. According to Boyle's law $V \propto \frac{1}{P}$ or $PV =$ constant.
6. When pressure is plotted against the reciprocal of volume ($1/V$), the curve obtained is a straight line passing through the origin.
7. A curve obtained by plotting pressure against volume at constant temperature is called isotherm which is rectangular hyperbola.
8. The relation between volume and temperature at constant pressure is given by Charles' law.
9. Charles' law states that the volume of a given mass of gas at constant pressure, increases or decreases 1 / 273 of the volume at 0°C for every degree centigrade rise or fall of temperature.
10. When the temperature is increased from 0°C to 273°C, the volume of a gas is doubled at constant pressure.
11. When the temperature is decreased to −273°C, the volume of a gas becomes zero.
12. The volume of a gas is directly proportional to the absolute scale of temperature at constant pressure. $V \propto T$ at constant pressure. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (or) $\frac{V_1}{V_2} = \frac{T_1}{T_2}$
13. The curve obtained by plotting volume against the temperature at constant pressure is called isobaric curve.
14. When the absolute temperature of a given mass of gas is doubled at constant pressure, its volume doubled.
15. Gay-Lussac's law: The relation between pressure and temperature at constant volume is given by this law.
16. Pressure of a given mass of gas at constant volume, increases or decreases by 1 / 273 of the pressure 0°C for every degree centigrade rise or fall of temperature.
17. At constant volume, the pressure of a gas becomes zero when the temperature is decreased to −273°C. At constant volume, when the temperature is increased from 0°C to 273°C the pressure is doubled. At constant volume, the pressure of a given mass of gas is directly proportional to the absolute temperature.
    $P \propto T$ at constant volume $P = KT$; $\frac{P}{T} = K$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ (or) $\frac{P_1}{P_2} = \frac{T_1}{T_2}$
18. When the absolute temperature of a given mass of gas is doubled at constant volume, the pressure is doubled.
19. The curve obtaining pressure against temperature at constant volume is called isochor or isometric curve.

20. Avogadro's Law: Equal volumes of gases under similar conditions of temperature and pressure contain equal number molecules. At constant temperature and pressure, the volume of a gas is directly proportional to the number of mole. $v \propto n$ at constant temperature and pressure $v = Kn$; $\frac{v}{n} = K$; $\frac{v_1}{n_1} = \frac{v_2}{n_2}$

21. Dalton law: At constant temperature, the total pressure of a gaseous mixture the sum of the partial pressures of the component gases provided there is no chemical reaction between the component gases. $P = p_A + p_B + p_C$

22. The pressure exerted by a component gas in the mixture of gases at a given temperature is called its partial pressure. The product of the total pressure and the mole fraction of a component gas gives the partial pressure. If the partial pressure of component A is p_A; then $p_A = X_A.P$ where X_A is mole fraction of A and P is the total pressure of the mixture.

23. Similarly the Partial pressure p_B of the component B is $p_B = X_B.P$

24. Dalton's law is not applicable to the following pairs of gases as they chemically react with each other.

    $NO + O_2$, $CO + Cl$, $H_2 + Cl$, $NH_3 + HCl$, $H_2 + F_2$ etc.

25. Graham's Law of diffusion: The property of intermixing of gases with each other is called diffusion. Gases diffuse more rapidly than liquids and solids. The diffusion of a gas depends on its density, vapour density and molecular weight. The volume of a gas diffused per second is called the rate of diffusion.

    Rate of diffusion $r = \frac{v}{t}$; where v = volume of gas diffused, t is time.

    Unit of rate of diffusion is c.c./sec or litre/sec.

26. The rate of diffusion of a gas is inversely proportional to the square root of its density or vapour density or molecular weight. Then if r_1 and r_2 are the rates of diffusion of two gases, d_1 and d_2 are their densities, $v.d_1$ and $v.d_2$ are their vapour densities and M_1 and M_2 are their molecular weights, then

    $$r \propto \frac{1}{\sqrt{d}};\ r \propto \frac{1}{\sqrt{vd}};\ r \propto \frac{1}{\sqrt{M}}$$

    $$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{vd_2}{vd_1}} = \sqrt{\frac{M_2}{M_1}}$$

    $$\frac{r_1}{r_2} = \frac{v_1}{t_1} \times \frac{t_2}{v_2} = \sqrt{\frac{M_2}{M_1}}$$

    If two gases diffuse the same volumes in different times then $\frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$

    Time taken for the diffusion of a given volume of a gas-is directly proportional to the square root of its molecular weight if two gases diffuse different volumes in the same time, then $\frac{v_1}{v_2} = \sqrt{\frac{M_2}{M_1}}$. The volume of a gas diffused is inversely proportional to the molecular weight of the gas.

27. Lighter gases diffuse faster than the heavier gases. Hydrogen being the lightest gas known as the highest rate of diffusion at a given temperature. Rate of diffusion of a gas is directly proportional to the r.m.s. velocity of the gas.
28. Gases having the same molecular weights have the same vapour densities and same rates of diffusion. Eg: CO_2, N_2O, C_3H_8 have the same rate of diffusion.
29. CO, N_2, C_2H_4 have the same molecular weight & have the same rates of diffusion.
30. The rates of diffusion of hydrogen and oxygen are in the ratio 4:1.
31. The rates of diffusion of He and SO_2 are in the ratio 4:1 & the rates of diffusion of methane and SO_2 are in the ratio 2:1.
32. The process of separation of a mixture of gases into components using their difference in the rate of diffusion is called atmolysis.
33. Atmolysis used in the separation of isotopes of uranium in the form of their hexafluorides, alarm in coal mines works on the principle of Graham's law of diffusion.
34. Diffusion of a gas through a narrow hole is called effusion.
35. Effusion also obeys Graham's law.
36. Ideal gas equation: Ideal gas equation is derived from Boyle's law, Charles' law and Avogadro's law. $PV = nRT$ is called ideal gas equation. R is called the universal or molar gas constant.
37. The value of R depends on the units of P and V. R = 0.0821 litre-atm / K /mole.
    = 82.1 ml-atm / K /mole = 62.4 lit-mm / K / mole
    = 8.314×10^7 erg / K /mole = 8.314 J / K / mole
    = 1.987 cal / K / mole = 0.002 kCal / K /mole
38. Gas constant for single molecule of gas is called Boltzmann constant (k). $k = \frac{R}{N}$
    $K = 1.38 \times 10^{-16}$ erg / K / molecule or 1.38×10^{-23} J / K /molecule.
39. A gas which obeys all gas laws under all conditions of temperature and pressure is called an ideal gas. Gases which deviate from ideal behaviour are called non-ideal gases or real gases. H_2 and He show nearly ideal behaviour. Real gases approach ideal behaviour at high temperature and low pressures. Real gases deviate from ideal behaviour at low temperatures and high pressures. No known gas is an ideal gas and all are real gases.
40. Kinetic theory of gases: Major contributors towards the development of kinetic theory of gases are Maxwell, Boltzmann, Clasius, Kronig etc. According to this theory
    (a) All gases are composed of extremely small particles called molecules.
    (b) Gas molecules are spherical and perfectly elastic.
    (c) Gas molecules constantly move with high velocities in all possible directions.
    (d) gas molecules continuously collide with one another and also the walls of the container.
    (e) the collisions made by the molecules on the walls of the container are responsible for the pressure of the gas.
    (f) the actual volume of the gas molecules is negligible when compared to the volume of the container.
    (g) there are no attractive or repulsive forces between the gas molecules.
    (h) there is no influence of gravitational force on the molecular motion.

(i) the average kinetic energy of the gas molecules is directly proportional to the absolute scale of temperature.

41. $PV = \frac{1}{3}mnc^2$ is the kinetic gas equation where c is the root mean square velocity of the gas.

42. The square root of the mean of the squares of the velocities of all molecules of a gas is called its root mean square velocity.

$$c = \sqrt{\frac{c_1^2 + c_2^2 + c_3^2 + c_n^2}{n}}$$

$$c = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{PV}{M}} = \sqrt{\frac{3P}{d}}$$

The root mean square velocity of a gas is directly proportional to the square root of the absolute scale of temperature.

43. When the molecules colloid with each other or, with the walls of the container, there will be change in molecular velocities.

44. The velocity of a molecule remains same for a very short period of less than 10^{-9} seconds.

45. The ratio of number of molecules with certain velocity to the total number of molecules always remain constant , at an instant.

46. The conclusions drawn from the above curves are

(a) The fraction of molecules having zero velocity is zero

(b) The fraction of molecules having either very low velocities (or) very high velocities is very small

(c) The peak of the curve x represents the most probable velocity.

(d) The velocity possessed by a maximum number of molecules of a gas is called the most probable velocity (Cp).

(e) At constant temperature, the order of molecular velocities is $C_p < \overline{C} < C$

(f) As the velocities of molecules increase the fraction of molecules having particular velocities also increases upto a maximum value and then decreases.

(g) When the temperature increases, the curve shifts to the right side. The height of the crest decreases and becomes flatten.

(h) Boltzmann equation is $n_1 = ne\text{-}^{Ei/KT}$.

Here n = total number of gas molecules

n_i = number of gas molecules possessing energy E_i.

E_i / KT = Boltzmann factor

K = Boltzmann constant,

T = Temperature in Kelvin scale.

(i) Gas molecules possess three types of velocities:

rms velocity (C): The square root of mean of squares of velocities of the gas molecules in the gas.

(j) A gas contains n molecules possessing the individual velocities C_1, C_2, C_3, --- C_n. Then the RMS velocity is $C = \sqrt{\frac{C_1^{\ 2} + C_2^{\ 2} + C_3^{\ 2} + \cdots C_n^{\ 2}}{n}}$

(k) When pressure and volume are given $C = \sqrt{\dfrac{3PV}{M}}$

(l) When P is expressed in dyne cm^{-2} and V in cm^3, the unit of C is cm/sec.

(m) When P is expressed in Newton m^{-2} and V in m^3, the unit of C is m/sec.

(n) When the density is given $C = \sqrt{\dfrac{3P}{d}}$.

(o) The gas with highest RMS velocity is H_2.

(p) RMS velocity can be calculated using the formula $C = \sqrt{\dfrac{3RT}{M}} = \sqrt{\dfrac{3PV}{M}} = \sqrt{\dfrac{3P}{d}}$

The ratio of RMS velocities of two gases is $\dfrac{C_1}{C_2} = \sqrt{\dfrac{T_1 M_2}{T_2 M_1}}$

(q) RMS velocity is used in the derivation of kinetic gas equation. Since it represents the velocity of all molecules of the gas.

(r) When the temperature is given $C = \sqrt{\dfrac{3RT}{M}} = 1.58 \times 10^4 \sqrt{\dfrac{T}{M}}$ cm / sec

(s) Average Velocity ($\bar{C}$):

Average velocity of gas molecules is the ratio of the velocities of gas molecules to the total number of molecules. It is denoted by ($\bar{C}$). If C_1, C_2, C_3, ----- C_n are the individual velocities of molecules, then the average velocity is $\vec{C} = \dfrac{C_1 + C_2 + C_3 + \cdots C_n}{n}$

(t) Average velocity can be calculated by the formula $\vec{C} = \sqrt{\dfrac{8RT}{\pi M}} = \sqrt{\dfrac{8PV}{\pi M}} = \sqrt{\dfrac{8P}{\pi d}}$. Average velocity $\bar{C} \equiv 0.9213 \times$ RMS velocity

Among three types of velocities Average velocity has least magnitude.

(u) Most Probable Velocity (Cp): Most probable velocity is the velocity possessed by maximum number of gas molecules. It is denoted by Cp. Most probable velocity can be calculated by the formula $C_p = \sqrt{\dfrac{2RT}{M}} = \sqrt{\dfrac{2PV}{M}} = \sqrt{\dfrac{2P}{M}}$.

Most probable velocity (Cp) = 0.8166 × RMS velocity. The magnitude of most probable velocity is in between RMS velocity and average velocity. The ratio of the molecular velocities is $C_P : \bar{C} : C = 1 : 1.128 : 1.224$

$\bar{C} = 1.128 \times C_p$, $C = 1.224 \times C_p$,

$C = 1.086 \times \bar{C}$, $\bar{C} = 0.921 \times C$, $C_p = 0.817 \times C$

At the same temperature the gas with lowest molecular weight possess the highest RMS or average or most probable velocities.

Non-ideal behavior of a Gas:

1. The factor $e\sim{}^{E_i/RT}$ is called the Boltzmann factor. It indicates the fraction of the molecules having an energy E or higher than E.
2. Real gases behave as ideal gases at very low pressure and high temperature. At high pressure and low temperature the deviations of real gases are more pronounced.

3. The assumptions of the kinetic theory are applicable to real gases only at low pressure $_v$ and high temperature. M
4. At high pressure and low temperature, the gases deviate considerably from the ideal behaviour.
5. Deviation can be expressed in terms of compressibility factor (z), expressed as
   $Z = \frac{PV_m}{nRT} = \frac{\text{Molar vol. of the gas}}{\text{nRT Molar vol. of perfect gas}}$
   In case of ideal gas, PV = nRT, z = 1
   In case of real gas, PV =£ nRT, $Z \neq 1$
6. Negative deviation: In such case, z < 1 gas is more compressible, due to attraction interaction.
7. Positive deviation : In such case, z > 1, gas is less compressible, due to repulsive interaction.
8. Nature of the gas, p ressure and temperature are the factors affecting the deviation.
9. Boyle's Temperature: Temperature at which a real gas exhibits ideal behaviour for considerable range of pressure is called Boyle's temperature.
10. Cause of deviation from the ideal behaviour are volume occupied by the gas molecules is negligible compared to the total volume of gas and there are no attractive forces between the gas molecules.
11. Vander Waals obtained the following equation for n moles of a gas =
    $$\left(\frac{P+an^2}{V^2}\right)(V-nb) = nRT$$
12. The constant 'b' in Vander Waals equation is known as excluded volume or co - volume. It is nearly four times the actual volume occupied by the gas molecules.
    Excluded or co - volume (b) = 4 x actual volume of gas molecules.
13. The attraction experienced by a given molecule is proportional to the concentration n/V of molecules in that container. As the attractions slow down the molecules, the molecules strike the waals less frequently and strike with a weaker impact. Reduction in pressure is proportional to the square of the molar concentration.
14. The greater the value of 'a', the greater the strength of Vander Waals forces and the greater is the ease with which a gas can be liquified.
15. Units of ‘ a’ are atm litre^2 mole^{-2} Units of ‘b’ are atm litre mole^{-1}
16. Exceptional behaviour of Hydrogen and Helium. Due to their small size, the attractive forces between the molecules are too small,
    $\frac{a}{V^2}$ is negligible then $\frac{PV}{RT} = 1 + \frac{Pb}{RT} \therefore z = 1$
17. Increase of pressure and,-"decrease of temperature tend to cause liquefaction of gases. The effect of temperature is, however more important.
18. Critical temperatureis defined as the temperature above which no gas can be liquefied how ever high the pressure may be. Critical temperature of CO_2 is 31.1 °C
19. At critical temperature the pressure needed to liquefy a gas is known as critical pressure.
20. The volume occupied by one mole of a gas at critical temperature and critical pressure is known as critical volume.
21. Relationship between critical constants and Vander Waal's constant.
    $V_c = 3b,\ T_c = \frac{8a}{27Rb}, P_c = \frac{a}{27b^2}$
22. Boiling point = 2/3 of critical temperatures.
23. When a gas under high pressure is allowed to expand into a region of low pressure, it suffers a fall in temperature. Cooling of gas by expansion from high pressure side to low

pressure is called joule- Thomoson effect.

24. To liquify gases like N_2 and O_2 the technique based on intermolecular forces is used. If the velocities of molecules are reduced to such lower values that neighbours can attract each other by their interaction or inter molecular attractions, then the cooled gas will condense to a liquid. The molecules are allowed to expand into available volume without supplying any heat from outside.
25. For hydrogen under normal condition repulsions dominated and $z > 1$. Therefore, it warms up in Joule - Thomson expansion.

    Inversion temperature $T_i = \frac{2a}{Rb}$
26. Enthalpy remains constant during Joule Thomson effect. At exact inversion temperature threre is no Joule Thomson effect above inversion temperature, there is heating.
27. Water exists either as solid ice or liquid water or gaseous steam, depending on its temperature.
28. Intermolecular forces are usually called as van der wall's forces.
29. Hydrogen bond in a specific intermolecular force limited to certain kinds of molecules,
30. Ion - dipole forces are mostly observed in aqueous solution of ionic substances in which dipole water molecules surround the ions.
31. Neutral polar molecules experience dipole dipole forces.
32. Dipole forces are attractive between unlike poles and repulsive between like poles and depend on orientation of the molecules.
33. Dipole forces are weak (energies of the order 3 - 4 kJ mol^{-1}) and significant only when the molecules are in close contact.
34. Dipole dipole interaction energy between stationary polar molecules as in solids in proportional to $1/r^3$ and that between rotating molecules is proportional to $1/r^6$ (Where 'r' is the distance between the polar molecules)
35. These forces exists among ionic and polar particles between individual atom or non - polar molecules.
36. An atom's (or) a molecule's electron cloud can be distorted by a nearby electric field, this property in known as polarisability.
37. A small molecule (or) atom in less polarisable and has smaller dispersion forces because it has few tightly held electrons.
38. Dipole - induced dipole forces are between polar molecules with permanent dipole moments and the molecules with no permanent dipole moments.
39. Permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming the electronic cloud.
40. The magnitude of these forces depends on : the magnitude of dipole moments of permanent dipole and polarisability of neutral molecule.

Liquid State:

- The intramolecular forces are stronger in liquids than gases
- Liquids have definite volume and are more denser than gases

Vapour Pressure: At constant temperature the presence exerted by the vapor over the liquid when it is in equilibrium is called vapour pressure.

→ At high altitude as the atmospheric presence is less, the liquids boil at low temperature.

- → The liquids has lowest energy when the surface area is minimum. Hence liquid droplets exist in spherical shape.
- → The regular gradation of velocity for layers in passing from one layer to the next layer is called "Laminar Flow".
- → As viscosity increases, liquids flow slowly.
- → Hydrogen bond and vandarwaals forces causes high viscosity
- → Glass is an extremely viscous liquid and properties resemble solids.
- → As the temperature increases, viscosity decreases since kinetic energy of molecules that overcome the intermolecular forces.
- → Mathematics expression of viscosity

$$F = \eta A \frac{dv}{dx} \qquad \left(n = Ae^{E/RT}\right)$$

η = Coefficient of viscosity
dv = Velocity gradient
dx = Distance between layers
A = Area of contact
F = Laminar flow

Question Bank:

1. The density of air in 0.001293gm per cc. Its vapour density will be:
   (1) 10 (2) 16
   (3) 1.43 (4) 14.3

2. Density of a gas x is eight times of that of a gas y. If M is the molecular weight of gas x, χ the molecular weight of gas y will be
   (1) 4.0M (2) 2.0N
   (3) 0.5M (4) 0.125M

3. Which of the following oxides of nitrogen is a coloured gas?
   (1) N_2O (2) NO
   (3) N_2O_4 (4) NO_2

4. The value of Vander Walls constant 'a' for the gasses O_2, N_2, NH_3 and CH_4 are 1.360, 1.3390, 4.170 and 2.253 atm.l^2/mol^2 respectively. The gas which can most easily be liquefied is:
   (1) O_2 (2) N_2
   (3) NH_3 (4) CH_4

5. The specific heat ratio r for inert gases is:
   1) 1.67 (2) 1.40
   (3) 1.33 (4) None

6. The gas with highest rate of diffusion among the following is:
   (1) NH_3 (2) O_2
   (3) N_2O (4) CO

7. Pressure Coefficient is:
   (1) 0.0367 (2) 0.00367
   (3) 0.367 (4) 3.67

8. If the rate of diffusion of H2 is 10 times greater than the gas x, the Molecular weight of X is:
   (1) 100 (2) 200
   (3) 50 (4) 1000

9. A vessel contain 1 gram of hydrogen 4 grams of Helium 7 grams of Nitrogen and 8 grams of oxygen. Which gas exerts more partial pressure?
(1) H_2 (2) He
(3) N_2 (4) O_2

10. If Grahams law is applied for Effusion, then it can be written as:
(1) $\frac{r_2}{r_1}=\sqrt{\frac{d_1}{d_2}}$
(2) $r_1r_2=\sqrt{d_1d_2}$
(3) $\sqrt{\frac{r_1}{r_2}}=\sqrt{\frac{d_2}{d_1}}$
(4) $r\propto\sqrt{d}$

11. At a particular temperature a vessel contain a mixture of 4.4 grams of CO_2 and 5.6 grams of N_2. If the total pressure of the gaseous mixture is 1.5 atm; the partial pressure of N_2:______
(1) 0.5atm (2) 1atm
(3) 0.25atm (4) None

12. The volume of 2 grams of a gas at 27°C and 380 mm pressure is 500 ml. The molecular weight of the gas is:
(1)197 (2) 200
(3) 19.7 (4) 190

13. 50 ml of O_2 diffuse in 30seconds from a porous pot. The time taken for the diffusion of the same volume of H_2 through that pot is:
(1) 60sec (2) 15sec
(3) 30sec (4) 7.5sec

14. The average kinetic energy of 16 grams of methane at –273°C is:
(1) 2.1×10^4ergs
(2) 0 ergs
(3) 0.21×10^4ergs
(4) 0.021×10^4ergs

15. The number of molecules present in 1litre of O_2 at STP is :
(1) $22.4\times6.023\times10^{23}$
(2) $\frac{22.4}{6.023}\times10^{23}$
(3) $\frac{6.023\times10^{23}}{22.4}$
(4) None of the above

16. A spherical balloon of 14cm, diameter is to be filled with H_2 at NTP from H_2 cylinder of volume 14.37 liters at STP then the number of the balloons can be filled with that H_2 gas present in the cylinder are:
(1) 300 (2) 3
(3) 10 (4) 100

17. The RMS velocity of N_2O at 30°C is equal to the RMS velocity of CO_2 at ____ temperature:
(1) 35°C (2) 40°C
(3) 30°C (4) 25°C

18. For a given gas, the product of pressure and volume is directly proportional to:
(1) To its mass
(2) Absolute temperature
(3) Critical temperature
(4) None of the above

19. Gases are completely missible. This is supported by the following law:

(1) Gay lussac law
(2) Graham's law of diffusion
(3) Avogadro's law
(4) None of the above

20. For a gaseous element the Cp/Cv indicates:

(1) Molecularity
(2) Atomicity
(3) Velocity
(4) Compressibility

21. The RMS velocity of O_2 at 20^0C is equal to the RMS velocity of SO_2 at a temperature of:

1) 313°C (2) 300°C
3) 293°C (4) 273°C

22. The value equivalent to Boltzman constant is:

(1) R / N (2) N / R
(3) R . N (4) R + N

23. Which of the following pair has same rate of diffusion?

(1) C_3H_8, CO_2 2) NO, CO
(3) NH_3, PH_3 (4) Cl_2, O_2

24. The volume occupied by 8.5grams of a gas with vapour density 8.5 at STP is:

(1) 5.6 liters (2) 11.2 liters
(3) 22.4 liters (4) 44.8 liters

25. **A:** 8 gm of CH_4 occupies 11.207 of volume at S.T.P
R: 1 mole of any gas at S.T.P occupies 22.4142

(1) **A** and **R** are true and **R** is the correct explanation of **A**
(2) **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false
(4) **A** is false, but **R** is true

26. **A:** At a given temperature 1 mole of all gases contain same Kinetic energy
R: Kinetic energy = (3 / 2) RT (for 1 mole of gas)

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

27. **A:** H_2 and O_2 have same R.M.S velocity at the same temperature
R: R.M.S velocity of a gas molecular is directly proportional to square root of is *T*.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not correct explanation of **A**
(3) **A** is true, but **R** is false
(4) **A** is false, but **R** is true

28. **A:** Poisonous gases can be diluted by the process of diffusion.
R: Ansil's alarm works on the principle of Graham law

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false
(4) **A** is false, but **R** is true

29. Assume that air is 21.0% O_2 and 79.0% N_2 by volume. If the barometric pressure is 740mm; the partial pressure of O_2 is closest to which one of the following?
(1) 155 mm (2) 310 mm
(3) 580 mm (4) 720 mm

30. A certain mass of dry gas at 27°C and 720 mm pressure has density 28. Its density at 7°C and 740mm will be:
(1) 30.80 (2) 38.0
(3) 30.08 (4) 83.0

31. When the temperature of 23 ml of dry CO_2 gas is changed from 10°C to 30°C at a constant pressure of 760 mm, volume of gas becomes closest to:
(1) 21.5 ml (2) 24.6 ml
(3) 25 ml (4) 23.0 ml

32. According to the Maxwell's Boltzmann distribution law for gases, the average transitional kinetic energy is:
(1) 0.5 KT (2) K.T
(3) 1.5 KT (4) 2.5 KT

33. The time taken for certain volume of gas to diffuse through a small hole is 2 minutes. Under the same conditions, an equal volume of O_2 took 5.65 minutes to pass through. The molecular weight of the gas is:
(1) 2.016 (2) 32
(3) 11.33 (4) 4

34. Two flasks *A* and *B* of 500ml each are respectively filled with oxygen and sulphur dioxide at 300°K and 1 atm pressure. The flasks will contain:
(1) The same number of atoms
(2) The same number of molecules
(3) The same amount of gases (in grams)
(4) More number of moles in flask *A* compared to flask *B*

35. 50 ml of H_2 diffuses from a vessel in 20 minutes. 45 ml of O_2 under similar conditions diffuse:
(1) 64 minutes (2) 72 minutes
(3) 20 minutes (4) 32 minutes

36. The average velocity of an ideal gas molecule at 27°C 0.3 m/sec. The average velocity at 927°c will be of that gas:
(1) 0.6 m/sec (2) 0.3 m/sec
(3) 1.2 m/sec (4) 3.0 m/sec

37. The weight of 11.2 liters of NH_3 at STP is:
(1) 7 gm (2) 8.5 gm
(3) 22.4 gm (4) 14 gm

38. Which of the following pairs of gases has some rate of diffusion at the given temperature?

(1) CO_2 and N_2O
(2) CO_2 and N_2O_4
(3) CO_2 and CO
(4) N_2O

39. The partial pressure of chlorine in a mixture of 71 grams of chlorine and 32 grams of SO_2 is_____ times to total pressure:
(1) $\frac{2}{3}$ (2) $\frac{3}{2}$ (3) $\frac{1}{2}$ 4) $\frac{1}{4}$

40. At constant temperature, when a gas is expanded, then the pressure:
(1) Increases
(2) Remains same
(3) Decreases
(4) Reduced to half

41. The impact on the wall by the gas molecules develops the
(1) Mass (2) Velocity
(3) Pressure (4) Temperature

42. At one atmosphere pressure and 0°C temperature the volume of one mole of a gas is equal to:
(1) One liter (2) 22.4 liters
(3) 11.2 liters (4) 1.2 liters

43. The volume of 10^{23}molecules of CO_2 at 1atm pressure and 273°C is:
(1)5 liters (2) 6 liters
(3) 7.5 liters (4)10.0 liters

44. When a gas expands, cooling occurs. This is because the molecular:
(1) Attraction decrease
(2) Attraction increase
(3) Remains same
(4) None of these

45. K.E of one mole of a gas at –73°C in Joules is:
(1) 2490 (2) 249.0
(3) 24900 (4) 24.9

46. Temperature at which a real gas obeys the ideal gas at a wide range of pressure is called:
(1) Boyle's Temperature
(2) Average Temperature
(3) Charles's Temperature
(4) Joule's Temperature

47. Which of the following pairs have the same rate of diffusion under similar conditions
a) N_2O, C_3H_8 b) C_2H_6, CO
c) CO_2, C_2H_2 d) N_2, CO
(1) b, c (2) a, d
(3) b, d (4) a, c

48. What is the density (in g lit^{-1}) of CO_2 at 400 K and exerting a pressure of 0.0821 atm (R = 0.0821 lit atm mol^{-1} k^{-1})
(1) 0.01 (2) 1.1
(3) 2.5 (4) 44

49. If the rms velocity of a gas at 100K is 10^4 cm sec^{-1}, what is the temperature (in 0°C) at which the rms velocity will be 3×10^4 cm sec^{-1}
1) 900 (2) 627
(3) 327 (4) 217

50. 4 grams of an ideal gas occupies 5.6035 liters of volume at 546 k and 2 atm, pressure. What is its molecular weight?
(1) 4 (2) 16 (3) 32 (4) 64

51. The kinetic energy of 4 moles of nitrogen gas at 127°C is ----- cals
(1) 4400 (2) 3200
(3) 4800 (4) 1524

52. At 27°C, one mole of an ideal gas exerted a pressure of 0.821 atmospheres. What is its volume in liters?
(R = 0.082 lit-atm / $mol^{-1}K^{-1}$)
(1) 300 (2) 30
(3) 0.3 (4) 3

53. A sample of gas has a volume of 0.2 lit measured at 1 atm pressure and 0°C. At the same pressure, but at 273°C, its volume will becomes
(1) 0.1lit (2) 0.8 lit
(3) 27.8 lit (4) 5.6lit

54. What is the value of gas constant R in J mol^{-1} K^{-1}?
(1) 82.1 (2) 8.31×10^{7}
(3) 8.314 (4) 0.0821

55. The volume of 10 moles of an ideal gas at 27°C and 1atm pressure is 1 lit. What is the volume of 20 moles of same gas at same pressure and temperature?
(1) 2 lit (2) 4 lit
3) 1 lit (4) 8 lit

56. The total pressure of a mixture of 6.4 grams of oxygen and 5.6 grams of nitrogen present in a 2 lit vessel is 1200 mm. What is the partial pressure of nitrogen in mm?
(1) 1200 (2) 600
(3) 900 (4) 200

57. NH_3 gas is liquefied more easily than N_2. Hence:
(1) Vander waal's constant a and b of $NH_3 >$ that of N_2
(2) Vander Waal's constant a and b of $NH_3 >$ that of N_2
(3) $a(NH_3) > a(N_2)$ but $b(NH_3) < b(N_2)$
(4) $a(NH_3) < a(N_2)$ but $b(NH_3) < b(N_2)$

58. Vender Waal's equation for one mole of CO_2 gas at low pressure will be :
(1) $\left(P+\frac{a}{V^2}\right)V = RT$
(2) $P(V-b) = RT-\frac{a}{V^2}$
(3) $P=\frac{RT}{V-b}$
(4) $P\left(\frac{RT}{V-b}-\frac{a}{V^2}\right)$

59. Compressibility factor for H2 behaving as real gas is
(1) 1
(2) $\left(1-\frac{a}{RTV}\right)$
(3) $\left(1+\frac{Pb}{RT}\right)$
(4) $\frac{RTV}{(1-a)}$

60. The deviation from the ideal gas behavior of the gas can be expressed as
(1) $Z=\frac{P}{VRT}$ (2) $Z=\frac{PV}{nRT}$
(3) $Z=\frac{nRT}{PV}$ (4) $Z=\frac{RT}{PT}$

61. The critical temperature of a substance is defined as:
(1) Temperature above which the substance decomposes

(2) The temperature above which a substance can exist only as a gas
(3) melting point of the substance
(4) boiling point of the substance

62. The Joule Thomson coefficient is zero at
(1) Inversion temperature
(2) Critical temperature
(3) Absolute temperature
(4) Below 0^0C

63. The behavior of bemporary gases like CO_2 approaches that of permanent gases like N_2, O_2 etc as we go
(1) Below critical temperature
(2) Above critical temperature
(3) Above absolute zero
(4) Beow absolute zero

64. Generally liquid drops assume spherical shape because:
(1) A spahere has maximum surface area
(2) A sphere has minimum surface area
(3) Sphere is symmetrical in shpape
(4) None of these

65. Find incorrect match
(1) Unit of surface energy = $J-m^{-2}$
(2) Unit of surface tension (T) = $N-m^{-1}$
(3) Molecules on the surface of liquid = less energy
(4) Minimum surface area of a liquest = Lowest energy state

66. Liquids show visocosity which is due to
(1) Creation of friction between the layers of the fluid.
(2) Inter molecular attraction forces of the liquid
(3) Inter molecular repulsion forces of the liquid
(4) None of these

67. Laminar flow of a liquid means
(1) Regular gradation of velocity for layers in passing from one layer to the next layer of a liquid
(2) Showing constancy in the velocity of layers of a liquid
(3) Increase in the velocity of layers from surface to bottom of a liquid
(4) All

68. Which has the maximum vesocosity?
(1) Water (2) Glycol
(3) Acetone (4) Ethanol

69. Theunit of viscosity is
(1) Poise (2) Millipoise
(3) Centipoise (3) All these

70. In SI system, the units of coefficient of viscosity, η are
(1) $Kg\ s^{-1}m^{-2}$ (2) $Kg\ m^{-1}s^{-1}$
(3) $Kg\ cm^{-1}s^{-1}$ (4) $g\ m^{-1}s^{-1}$

71. London forces arises
(1) Due to instaneous dipole on one atom of a molecule

(2) Due to permanent dipole on one atom of a molecule
(3) Due to the presence dipole on one atom of a molecule
(4) None

72. F_2 is gas but I_2 is solid, because
(1) Larger London forces are present in I_2 when compared with F_2
(2) Lesser number of london forces are present in I_2 when compared with F_2
(3) F_2 and I_2 has same extent of london forces
(4) I_2 has low bond dissociation energy

73. The inter molecular forces present in inert gases are
(1) Ion – ion
(2) Ion – dipole
(3) Dipole – dipole
(4) Dispersion

KEY

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

5. STOICHIOMETRY

Synopsis

Significant Figures:

1. Every experimental measurement has some amount of uncertainty associated with it. However, one would always like the results to be precise and accurate. Precision and accuracy are often "rferred to while we talk about the measurement.
2. ***Precision*** refers to the closeness of various measurements for the same quantity. However, ***accuracy*** is the agreement of a particular value of the result. For example, if the true value for a result 2.00 g and a student 'A' takes two measurements and reports the results as 1.95g and 1.93g. These values are precise as they are close to each other but are not accurate. Another student repeats the experiment and obtains 1.94g and 2.05g as the results for the measurements. These observations are neither precise nor accurate. When a third student repeats these measurements and reports 2.01 g and 1.99g as the result. These values are both precise and accurate.
3. The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures. ***Significant figures*** are meaningful digits which are known with certainty. The uncertainty is indicated by writing the certain digits and the last uncertain digit. Thus, if we write a result as 11.2 mL, we say the 11 is certain and 2 is uncertain and the uncertainty would be ± 1 in the last digit is always understood.
4. The quantitative relationship existing between the quantities of reactant and products in a chemical reaction is known as stoichiometry.
5. The short form representation of chemical reaction using formulae of reactant and product is known as chemical equation.
6. The balanced chemical equation is called a stoichiometric equation.
7. The numerical placed before the formulae of reactant or product in order to balance the chemical equation is known as stoichiometric co-efficient.
8. The number of atoms present in a molecule is called its atomicity. Inert gases are monatomic. Hydrogen, Oxygen, Nitrogen and halogens are diatomic Sulphur is octa-atomic (S_8) and white phosphorus's tetra atomic (P_4) Arsenic and antimony are tetra atomic.
9. One a.m.u. is also known as one Dalton or one Aston.
10. Atomic mass of an element is the average of the isotopic masses (in a.m.u.) of the isotopes present in it.

$$\text{Number of moles } (n) = \frac{\text{Mass in grams} (G)}{\text{Gram Molecular Weight } (M)}$$

One mole of any substance (or one mole of a mixture of substances) contains the same number of molecules namely 6.023×10^{23} molecules. This number is known as Avogadro number (N).

11. One mole of any gas or vapour (or mixture of gases) at STP occupies a volume of 22.4 litres. This is known as gram molar volume. (G.M.V)

One mole = Numerical value of the molecular weight of the substance expressed in grams = mass of 6.023×10^{23} molecules of the substance = mass of 22.4 litres of the gas or vapour at STP.

12. The numerical value of the atomic weight of an element expressed in grams is known as a gram atomic weight or a gram-atom of that element.
E.g; One gram-atom of carbon = 12 grams of carbon.

13. $$\text{Number of gram - atoms } (n) = \frac{\text{Mass in grams of the element } (g)}{\text{Atomic Weight of the element}}$$

14. One gram ion or a mole of ions contains 6.023×10^{23} ions.

15. Numerical value of the formula weight of a substance expressed in grams is called a gram formula weight contains 6.023×10^{23} particles having that formula.

16. The value of 1 a.m.u. in gram is 1.667×10^{-24} gram. This is known as avogram. 1 a.m.u. = 1.667×10^{-27} kg.

17. Vapour density of a gas or vapour = $\dfrac{\text{Density of the gas}}{\text{Density of Hydrogen}}$ (Under the same conditions)

18. Molecular weight = density of the gas at STP × 22.4 lt

19. Vapour density of a gas = density of the gas at STP × 11.2 lt 2 × vapour density = Molecular weight.

20. In the case of a mono-atomic element one mole = one gram-atom. For example one mole of silver = one gram-atom of silver.

21. Molecular weight
= density of the gas at STP in g/L x 22.4

22. Vapour density of a gas = density of the g at STP x 11.2

23. Molecular weight is twice the vapour densi of a substance.

24. In a balanced chemical equation, always tv substances are in 1:1 ratio of their equivalents.

25. The numerical magnitude of the equivaler. weight of a substance need not necessarly a fixed value.

26. Equivalent weight of an acid

$$E_{Acid} = \frac{\text{Formula weight}}{\text{Basicity}}$$

27. Equivalent weight of base

$$E_{Base} = \frac{\text{Formula weight}}{\text{Acidity of base}}$$

28. Equivalent weight of salt

$$E_{Salt} = \frac{\text{Formula weight of the salt}}{\text{Total charge of the cation or anion of the salt}}$$

29. Equivalent weight of an element

$$\frac{\text{Atomic weight}}{\text{Valency of the element}}$$

30. $$E_{Fe^{+2}} = \frac{56}{2} = 28$$

31. According to modern electronic concept oxidation is a process in which electrons are removed from an atom or ion. Hence de electro nation is oxidation. When the oxidation number of an atom increases it is said to be oxidized.
32. Oxidation and reduction is a single reaction in which one substance is oxidized and another substance is reduced. It is called a redox reaction.
33. A reaction in which the same species undergoes both oxidation and reduction is called a disproportionate reaction.
34. The charged state to which an atom reaches after getting oxidized or reduced is called its oxidation number or oxidation state. Oxidation state indicates the state of chemical combination of the atom.
35. The oxidation number of an atom in a molecule is the formal charge [imaginary charge] which the atom would acquire if the bond pairs of electrons are allotted to the more electronegative atom.
36. The oxidation number of an atom in the element state is zero.
37. The oxidation number of a mono atomic ion is equal to its charge.
38. The sum of the oxidation numbers of all the atoms in a molecule is zero.
39. The sum of the oxidation numbers of all the atoms in a polyatomic ion is equal to its charge.
40. Fluorine always exhibits a −1 oxidation number in its compounds because it is the most electronegative element and it can gain only one electron.
41. Alkali metals always exhibit a +1 oxidation number in their compounds.
42. Alkaline earth metals always exhibit a +2 oxidation number in their compounds.
43. Oxygen generally exhibits an oxidation number −2 except in peroxides (−1) in super oxides (−1/2) and in ozonides (−1/3). When oxygen combines with fluorine, it exhibits a positive oxidation number, in O_2F_2 the oxidation number of oxygen is +1 and in OF_2 the oxidation number of oxygen is +2.
44. The oxidation number of Hydrogen is generally +1 except when it combines with metals to form hydrides in which its oxidation number is −1. [H^-].
45. The oxidation number of any representative element never exceeds its group number in the periodic table. The maximum oxidation number of any representative element is equal to its group number except in the case of oxygen and fluorine.
46. Calculation of Oxidation number: Suppose we require the oxidation number of Manganese in Potassium permanganate [$KMnO_4$]. Oxidation number of potassium is +1. Oxidation number of Manganese is x oxidation number of each oxygen is −2.
47. Common oxidizing agents and reducing agents: An element present in its lower oxidation state in a compound can act as a reducing agent. Eg., H_2S,SO_2, $Na_2S_2O_3$, $FeSO_4$, $SnCl_2$. An element present in its highest oxidation state in a compound can generally act as an oxidizing agent. Eg.,$KMnO_4$, $K_2Cr_2O_7$, HNO_3,Conc.H_2SO_4, $HClO_4$
48. A chemical equation is a short scientific representation of a chemical reaction.
49. A balanced chemical equation obeys law of conservation of mass.
50. The following steps are involved in balancing a redox reaction equation by the oxidation number method.

51. Note the oxidation number of the all the atoms on either side and identify the atoms whose oxidation numbers have changed.
52. Generally the oxidation number of one atom decreases and that of another atom increases.
53. Note the change in the oxidation numbers.
54. Then balance all atoms except hydrogen and oxygen, finally balance hydrogen and oxygen.
55. Balancing of redox reaction equations by the half-cell reaction method or ion-election method.
56. Represent first the ionic equation.
57. Indicate the oxidation half-reaction and reduction half reaction separately.
58. Balance the half reaction separately. While balancing the half reactions balance the atoms first. In acid medium include number of water molecules where there is deficiency of oxygen and include enough number of H^+ ions on the side where there is a deficiency of Hydrogen.
59. In alkaline medium include enough number of OH ions on the side where there is a deficiency of oxygen and enough number of water molecules where there is a deficiency of Hydrogen. After balancing the atoms, balance the charge.
60. To balance the charge includes enough number of electrons wherever necessary.
61. Multiply the reduction half reaction with the number of electrons lost in oxidation and multiplies the oxidation half reaction with the number of electrons gained in reduction.
62. Add the two half reactions after canceling the electrons in the final balanced equation not only the atoms but the net charge also should be balanced.
63. Empirical and molecular formulae:

    Percentage composition:- The masses of different elements present in 100g. of a compound is called the percentage composition of the compound.
64. The formula which shows the simple ratio between the atoms of the an element in a molecule of the substance is called empirical formula.
65. The formula that shows the exact number of atoms of each element in a molecule of the substance is called molecular formula.
66. Molecular weight of a substance can be determined by different methods like.
    a. Vapour density method
    b. Depression of freezing point method
    c. Elevation of boiling point method.

    Molecular formula = n x Emperical formula.

    Where $n = \frac{\text{molecular weight}}{\text{empirical formula}}$.

    Molecular weight = 2 $\times$ Vapour density.
67. The molecular weight of volatile compounds are determined by Victor Mayer's method.
68. Decomposition reactions are the opposite of combination reactions.

    eg: $2HgO \rightarrow 2Hg + O_2$
69. In the displacement reactions the place of one species in its compound is taken up by other species

    Eg: $Zn_{(s)} + CusO_{4(aq)} \rightarrow ZnSO_4 + Cu$

70. Disproportionation reactions involve the same element in the given form to under go both oxidation and reduction simultaneously.

71. The inverse of disproportionation is comproportionation. In comproportionat ion' reactions, two species with the same element in two different oxidation states form a single product in which the element is in an intermediate oxidation state.

72. In titrimetric analysis the substance of known concentration is called the titrant and the substance being titrated is called the titrand.

73. The point at which the titrand just completely reacts is called the equivalence point or the theoretical point or stoichiometric end point.

74. A compound added to reacting solutions that undergo an abrupt change in a physical property usually colour is called indicator.

75. The species participating in oxidation and reduction half reactions is called redox couple.

76. The solutions in the two beakers in a Galvanic cell are connected by a salt bridge. Salt bridge is a U tube containing an inert electrolyte solution like KC*l* made in the form of semisolid in agar-agar.

77. Standard electrode potential for hydrogen electrode is zero.

78. If the standard reduction potential is negative then it is a stronger reducing agent than H^+ / H_2 couple. If standard reduction potential positive then it is a weaker reducing agent than the H^+ / H_2 couple

Question bank:

1. What is the volume (in litres) of CO_2 liberated at STP, when 2.12 grams of sodium carbonate (MW=106) is treated with excess dilute HCl?
(1) 2.28 (2) 0.448
(3) 44.8 (4) 22.4

2. Which of the following reactions does not involve the change in oxidation state of metal?
(1) $VO^{-2} \to V_2O_3$
(2) $Na \to Na^+$
(3) $CrO_4^{2-} \to Cr_2O_7^{2-}$
(4) $Zn^{2+} \to Zn$

3. What are the oxidation numbers of 'N' in NH_4NO_3?
(1) +3, −5 (2) −3, +5
(3) +3, +6 (4) −2, +2

4. What is the volume (in litres) of oxygen at STP required for complete combustion of 32 g of CH_4 (mol wt of CH_4 = 16)
(1) 44.8 (2) 89.6
(3) 22.4 (4) 179.2

5. How many moles of acidified $FeSO_4$ can be completely Oxidised by one mole of $KmnO_4$
(1) 10 (2) 5 (3) 6 (4) 2

6. 50 g of calcium carbonate was completely burnt in air. What is the weight (in grams) of the residue?
(1) 2.8 (2) 28
(3) 4.4 (4) 44

7. 7.5 g of a gas occupy 5.6 litres as STP. The gas is
(1) NO (2) N_2O
(3) CO (4) CO_2

8. 4 grams of hydrocarbon complete combustion gave 12.571 grams

of CO_2 and 5.143 grams of water. What is the empirical formula of the hydrocarbon?

(1) CH (2) CH_2
(3) CH_3 (4) C_2H_3

9. In a balanced equation H_2SO_4 + xHI → H_2S + yI_2 + zH_2O. The values of x; y and z are

(1) $x = 3; y = 5; z = 2$
(2) $x = 4; y = 8; z = 5$
(3) $x = 8; y = 4; z = 4$
(4) $x = 5; y = 3; z = 4$

10. In acidic medium Dichromate ion oxidises Ferrous ion to Ferric ion. If the grammolecular weight of potassium dichromate is 294gm, its equivalent weight is

(1) 294 (2) 147
(3) 49 (4) 24.5

11. What is the volume (lit) of oxygen required at STP to completely convert 1.5 moles of sulphur into sulphur dioxide

(1) 11.2 (2) 22.4
(3) 33.6 (4) 44.8

12. An organic compound containing C and H has 92.3% carbon. Its empirical formula is

(1) CH (2) CH_3
(3) CH_2 (4) CH_4

13. 10 g of $CaCO_3$ is completely decomposed to X and CaO. X is passed into an aqueous solution containing one mole of sodium carbonate. What is the number of moles of sodium bicarbonate formed?

(1) 0.2 (2) 0.1
(3) 0.01 (4) 10

14. The percentage of water in the vitriol ($CuSO_4$, $5H_2O$) is

(1) 18% (2) 36%
(3) 64% (4) 75%

15. A sample of an alloy of aluminium contains 0.0972 mol Al and 0.0381 mol Mg. What are the mass percentage of Al and Mg in the alloy

(1) 74 Al, 26 Mg
(2) 26 Al, 74 Mg
(3) 15 Al, 85 Mg
(4) 85 Al, 15 Mg

16. A certain compound containing only carbon and oxygen has an approximate molecular weight of 290. On analysis it is found to contain 50% by weight of each element. Then the molecular formula of the compound is

(1) C_4O_3 (2) C_2O_3
(3) $C_{12}O_9$ (4) C_2H_8

17. The ratio of the molecular weights of two gases A and B is 1:2. Then the ratio of their vapour densities is

(1) 1 : 2 (2) 1 : 3
(3) 2 : 1 (4) 1 : 1

18. If the analysis of hydrocarbon indicated 85.6% carbon, the empirical formula the compound is

(1) C_3H_6 (2) C_2H_4
(3) C_6H_{12} (4) CH_2

19. The empirical formula of a compound is CH_2O. Its vapour density is 30 / 2 Its molecular formula is

(1) CH_2O (2) $C_2H_4O_2$
(3) $C_3H_6O_3$ (4) CH_3OH

20. The weight of 11.2 litre of any gas at STP represents its

(1) vapour density

(2) gram molecular weight
(3) gram atomic weight
(4) gram equivalent weight

21. The vapour density of gas *A* is four times that of *B*. If molecular mass of *B* is *M*, then molecular mass of *A* is
(1) *M* (2) 4*M*
(3) *M* / 4 (4) 2*M*

22. On analysis a certain compound was found to contain iodine and oxygen in the ratio of 254 g of iodine (at mass 127) and 80 g oxygen (at mass 16) What is the formula of the compound
(1) IO (2) I_2O
(3) I_5O_3 (4) I_2O_5

23. 60 g of a compound on analysis gave 24 g C, 4 g H and 32 g O. The empirical formula of the compound is
(1) $C_2H_4O_2$ (2) $C_2H_2O_2$
(3) CH_2O_2 (4) CH_2O

24. Mass of 6.022×10^{23} electrons is
(a) 0.55 mg (b) 55 mg
(c) 5.5×10^{-4} g (d) 9.8×10^{-31} g
(1) a and b (2) d
(c) a and c (4) all of right

25. 142 g of chlorine represents
(a) 2 g molecules of chlorine
(b) 4 moles of chlorine atoms
(c) 2 moles of Cl atoms
(d) 6.02×10^{23} Cl_2 molecules
(1) a and b (2) a, b & c
(3) b, c & d (4) all are right

26. In the reaction $4KI + 2CuSO_4 \rightarrow 2CuI + I_2 + 2K_2SO_4$. The equivalent weight of Cu is
(1) equal to its atomic weight
(2) double its atomic weight
(3) one-half of its atomic weight
(4) four times to its atomic weight

27. Positive oxidation number of an element indicates that it
(1) is an element form
(2) has been oxidized from elemental state
(3) has been reduced from elemental state
(4) does not change

28. A balanced equation will be in accordance with the law of
(1) multiple proportion
(2) conservation of mass
(3) reciprocal proportion
(4) both (1) and (3)

29. The number of carbon atoms present in 2.8 grams of carbon monoxide are
(1) 3.01×10^{23} (2) 3.01×10^{24}
(3) 6.02×10^{23} (4) 6.02×10^{22}

30. Maximum number of electrons are present in
(1) 2.24 litres of SO_2 at STP
(2) 0.2 moles of NH_3
(3) 1.5 gram moles of oxygen
(4) 2 mole atom of sulphur

31. The gas having same number of molecules as 16 g of oxygen is
(1) 16 g of O_3 (2) 16 g of SO_2
(3) 48 g of SO_3 (4) 1 g of H

32. Equal amounts of Zn is allowed to react with excess of H_2SO_4 and excess of NaOH separately. The ratio between the weights of H_2 formed is
(1) 1 : 2 (2) 1 : 1
(3) 1 : 3 (4) 2 : 1

33. The total number of species present in 1 mole of potash alum

in terms of avagardo number N are

(1) $3N$ (2) $5N$

(3) $8N$ (4) $32N$

34. If 0.7 moles of $BaCl_2$ is heated with 0.4 mole of potassium sulphate, number of moles at barium sulphate formed are

(1) 0.7 (2) 0.4

(3) 0.35 (4) 0.2

35. A gaseous mixture contains oxygen and nitrogen in 1 : 4 ratio by weight. The ratio of their number of molecules is

(1) 1 : 4 (2) 1: 8

(3) 7 : 32 (4) 3 : 16

36. A compound of phosphorus and sulphur contains 1.24 gm of phosphorus in 2.2 g of the compound. The formula of the compound is

(1) P_4S_3 (2) P_3S_4

(3) P_3S_2 (4) P_2S_3

37. Which of the following reactions does not involve change in the oxidation state of metal

(1) $VO^{2+} \rightarrow V_2O_3$

(2) $Na \rightarrow Na^+$

(3) $CrO_4^{2-} \rightarrow Cr_2O_7^{2-}$

(4) $Zn^{2+} \rightarrow Zn$

38. An aqueous solution of 6.3 g oxalic acid dilydrate is made upto 250 ml. The volume of 0.1N NaOH required to completely neutralize 10 ml of this solution is

(1) 40 ml (2) 20 ml

(3) 10 ml (4) 4 ml

39. How many milliliters of 1M H_2SO_4 solution will be neutralized by 10 ml of 1 M NaOH solution.

(1) 10 (2) 20

(3) 2.5 (4) 5

40. 4 g of a hydrocarbon on complete combustion give 12.571 g of CO_2 and 5.143 g of H_2O. What is the empirical formula of the hydrocarbon?

(1) CH (2) CH_2

(3) CH_3 (4) C_2H_3

41. The molecular formula of a compound which contains 30.5% of N_2, 69.5% of O_2 by weight and which has a molecular weight of 92 is

(1) NO (2) N_2O

(3) NO_2 (4) N_2O_4

42. A mixture of ethane and neon having a volume of 20 ml requires 49 ml of O_2 for complete combustion. The percentage of ethane in the mixture is

(1) 70% (2) 30%

(3) 35% (4) 64%

43. **A:** Concentrate NaOH reacts with fluorine to liberate oxygen gas.

R: In the above reaction oxygen undergoes disproportion.

The correct answer is

(5) Both **A** and **R** are true and **R** is the correct explanation of **A**

(6) Both **A** and **R** are true and **R** is not the correct explanation of **A**

(7) **A** is true, but **R** is false

(8) **A** is false, but **R** is true

44. **A:** 2 g of hydrogen contains Avogardo number of molecules.

R: One mole of ideal gas at STP occupies 22.4 litres.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false
(4) **A** is false, but **R** is true

45. **A:** In bleaching powder average oxidation number of chlorine is zero.

R: Oxidation number of chlorine is always zero.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false

46. **A:** The reaction $NH_3 + O_2 \rightarrow NO + H_2O$ cannot be easily balanced by ion election method.

R: Ion-election method is not suitable for vedox reaction take place in molecular state.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false

47. **A:** In the reaction $MnO_4^- + 8H^+ + Xe^- \rightarrow Mn^{2+} + 4H^2O$. The value of X = 5.

R: The oxidation state of Mn changes from +1 in MnO_4^- to +2 in Mn^{2+}.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true, but **R** is false

48. Match the following.

List – I ***Compound***	**List – II** ***Oxidation state of H^+***
A) CaH_2	1) –1
B) H_2O_2	2) +1
C) NaH	3) +2
D) H_2O	4) – ½
	5) + ½

The correct matching from List–I and List–II is

	A	**B**	**C**	**D**
(1)	1	2	1	2
(2)	1	2	3	5
(3)	5	2	3	4
(4)	4	3	2	5

49. Match the following.

List – I ***Half Reaction***	**List – II** ***No. of H^+ ions involved***
A) $Cr(OH)_3 \rightarrow Cr^{3+}$	1) $8H^+$
B) $CH^- \rightarrow CNO^-$	2) $2H^+$
C) $NH_2^- \rightarrow NO_2^-$	3) $3H^+$
D) $MnO_4^- \rightarrow Mn^{2+}$	4) $4H^+$
	5) $6H^+$

The correct matching from List–I and List–II is

	A	**B**	**C**	**D**
(1)	1	2	5	4
(2)	3	5	1	2
(3)	3	2	5	1
(4)	2	4	3	2

50. Match the following.

List – I ***Compound***	**List – II** ***Oxidation state of Cl***
A) HClO	1) +1
B) $HClO_4$	2) –1 & +1
C) $CaOCl_2$	3) –1 & 0
D) $HClO_3$	4) +7
E) Cl_2O	5) +5

The correct matching from List–I and List–II is

	A	**B**	**C**	**D**	**E**
(1)	1	2	3	4	5
(2)	5	4	3	2	1
(3)	1	4	2	5	1
(4)	2	5	1	3	4

KEY

1) **2**	2) **2**	3) **2**	4) **3**	5) **2**	6) **2**	7) **1**	8) **2**	9) **3**	10) **3**
11) **3**	12) **1**	13) **1**	14) **2**	15) **1**	16) **3**	17) **1**	18) **4**	19) **1**	20) **1**
21) **2**	22) **4**	23) **4**	24) **2**	25) **4**	26) **1**	27) **2**	28) **2**	29) **4**	30) **4**
31) **4**	32) **2**	33) **4**	34) **2**	35) **3**	36) **1**	37) **3**	38) **1**	39) **4**	40) **2**
41) **4**	42) **1**	43) **3**	44) **2**	45) **3**	46) **1**	47) **3**	48) **1**	49) **3**	50) **3**

In which of the following numbers all zeros are significant?

) 0 00004 2) 0.0060 3) 20 000 4) 0 800

2) On dividing 0.46 by 15 374 the actual answer is 0 0292 6 The correctly reported answer will be

1) 0 02 2) 0 029 3) 0 0292 4) 0.02924

3) The actual product of 4 327 and 2.8 is 12 1156 The correctly reported answer will be

1) 12 2) 12.1 3) 12 12 4) 12.116

4) After rounding off 1..235 and 1 225 we will have their answers respectively as

1) 1 23 1.22 2) 1.24 1.123 3)1.23 1.23 4)1 24 1 22

5) The correctly reported difference of 16.4215 and 6 01 will have significant figures equal to

1) three 2) four 3) five 4) six

6) The answer of the calculation $\frac{2.568 \times 5.8}{4.168}$ in significant figures will be:

1) 3 579 2) 3 570 3) 3.57 4) 3.6

7) The correctly reported answer of the addition of 154.2 , 6 142 and 23 will be

1) 183 352 2) 83.35 3) 183.4 4) 183

8) The correctly reported answer of the addition of 4.523, 2.3 and 6.24 will have significant figures

1) two 2) three 3) four 4) five

9) The number of significant figures in $N_0 = 6.022 \times 10^{23}$ i e. Avogadro s number are

1) three 2) four 3) five 4) can be any of these

10) The number of significant figures in 10500 are

1) three 2) four 3) five 4) can be any of these

11) The number of significant figures in 0.0645 are

1) two 2) three 3) four 4) five

12) Irrespective of the source pure sample of water always yields 88 89% mass of oxygen and 11.11% mass of hydrogen This is explained by the law of

1) Conservation of mass 2) Constant proportions

3) Multiple proportions 4) Constant volume

13) The mass of nitrogen per gram in hydrazine is exactly one and half the mass of nitrogen in the compound ammonia. The fact illustrates the

1) Law of conservation of mass 2) Multiple valency of nitrogen

3) Law of multiple proportion 4) Law of definite proportion.

14) The % of copper and oxygen in samples of CuO obtained by different methods were found to be the same This proves the law of

1) Constant Proportion 2) Reciprocal Proportion

3) Multiple Proportion 4) Conservation of mass

15) Which of the following pairs can be cited as an example to illustrate the law of multiple proportion?

1) NO_2, K_2O 2) CaCO MgO 3) Al_2O_3, Cr_2O_3 4) CO CO_2

16) The law of conservation of mass holds good for all of the following except.

1) all chemical reactions 2) Nuclear reaction
3) Endo thermic reactions 4) Exo thermic reactions

17) 10 L of N and 30 L of H_2 combine to form ____________ litres of NH

1) 20 2) 40 3) 30 4) 10

18) Which of the following is the best example of law of conservation of mass.

1) 12 gm of carbon combines with 32 gm of oxygen to form 44 gm of CO
2) When 72 gm of carbon is heated in vaccum and no change n mass takes place.
3) The weight of a piece of platinum is the same before and after heating in air.
4) 2 grm of Hydrogen combines with 32 gm of

19) One pa of an element (A) combines with t o parts of nother element (B). 6 parts of element (c) combines with 4 parts of (B) If d (c) combine together the ratio of their ights will be governed by

law of definite proportion of multiple proportion
reciprocal proportion a f conservation of mass

20) Law of multiple proportions is illustrated by one of the o lowing pairs

S and $S0_2$ 2) NH_3 and NO
a S and Na_3P 4) N_2 and NO

21) In th presence of dilute sulphuric acid the qui lent eight ofpotassiurn perrnanganate

) 1 5 of its molecular eight) 1 6 o it molecular eight
) 1 10 of its molecular weight 4) 1/2 of its melecular weight

22) 2g. of a divalent metal combines with 0.16g. of oxygen. Molecular weight of its oxide is

1) 200 2) 108 3) 100 4) 216

23) 2.0 g. of a metal oxide on strong heating gave 112ml.of O_2 gas at STP. Equivalent weight of the metal is

1) 112 2) 56 3) 50 4) 92

24) The equivalent weight of an element is 4. Its chloride has a vapour density 59.25 Then the valency of the element is

1) 4 2) 3 3) 2 4) 1

25) If 1.2g of metal displaces 1.12 litre of hydrogen at normal temperature and pressure

equivalent weight of the metal would be

1) 12 2) 24 3) 1.2 ‘11.2 4) 1.2 + 11.2

26) The equivalent mass of an element is 3. The vapour density of its volatile chloride is 77. The atomic weight ofthat element is

1)3 2)6 3)9 4)12

27) 0.23 g. of a metal when put into a jar ore chlorine produced 0.585g of its chloride The equivalent mass of the metal is

1) 21 2) 23 3) 26 4) 27

28) 9g. of trivalent metal combines with 8g. of oxygen. The atomic mass of the metal is

1)9 2)18 3)27 4)36

29) The equivalent weight of $K_2Cr_2O_7$ as an oxidising agent in acid medium is

1) 2 3 of its mol.wt. 2) 1 3 of its mol.wt.

3) 1 6 of its mol.wt. 4) 1 2 of its mol.wt.

30) If the equivalent weight of a trivalent metal is 32.7, the molecular weight of its chloride is

1) 68.2 2) 103.7 3) 204.6 4) 32 7

31) Equivalent weight of Pyrophosphoric acid is ($H_4P_2O_7$)

1 M.W/1 2 M.W/2 3) M W/4 4) M.W/3

32) Equivalent weight of hydrated oxalic acid is

1) 44 2) 45 3 63 4) 126

33) An organic compound has an empirical formula CH_2O Its V.D. is 45. The molecular formula of compound is:

1) CHP 2) C_2HP 3) C_2HP_2 4) C_3HP_3

34) If 75.2 % of compound is carbon and the rest of the weight is hydrogen, the formula of the compound is

1) C_3H_6 2) C_2H_2 3) C_2H_4 4) CH_4

35) An organic compound contains 40 % C, 6.66%H. The empirical formula of the compound is

1) CH_2 2) CH_2O 3) CHO 4) CHO_2

36) If the molecular weight of the compound is CH_2O. If the molecular weight of the compound is 180, the molecular formula is

1) $C_6H_{12}O_6$ 2) $C_3H_6O_3$ 3) $C_4H_8O_2$ 4) $C_5H_{10}O_5$

37) An organic compound on analysis gave C = 39.9%, H = 6.7 % and 0=53.4%. The empirical formula of the compound is

1) CHO_2 2) CH_2O 3) $C_2H_2O_2$ 4) CHO

38) The empirical formula of a compound is CH2. One mole of this compound has a mass of 42 gm. Its molecular formula is :

1) CH_2 2) C_2H_2 3) C_3H_6 4) C_3H_8

39) Equivalent weight of metal ‘M’is 12. Equivalent weight of Y in the compound MY_2 is (M.W of MY_2 is 36)

1) 24 2) 6 3) 12 4) 16

40) The empirical formula of a compound is CH_2O. If 0.0835 moles of the compound contains 1 gm of hydrogen, the molecular formula of the compound is

1) $C_2H_4O_2$ 2) $C_3H_6O_3$ 3) $C_6H_{12}O_6$ 4) $C_4H_8O_4$

41) On analysis, a certain compound was found to contain 254 grams of Iodine and 80 grams of oxygen. The formula of the compound is

1) IO 2) I_2O 3) I_5O_3 4) I_2O_5

42) One gram of a certain compound has 0.2729 gram of carbon and 0.7271 grams of oxygen. The formula of the compound is

1) CO 2) CO_2 3) C_3O_2 4) C_4O_2

43) 3.0 gms of an organic compound on combustion gives 8.8 gm of CO_2 and 5.4gm of water. The empirical formula of the compound is

1) CH_3 2) C_2H_4 3) C_2H_2 4) C_2H_6

44) An organic compound having molecular mass 60 .s found to contain C = 20%, H = 6.67% and N = 46.67% while rest is oxygen. On heating it gives NH_3 along with a solid residue. The solid residuegive violet colour with alkaline copper sulphate solution. the compound is

1) CH_3NCO 2) CH_3CONH_2 3) $(NH_2)_2$ CO 4) $CH_3CH_2CONH_2$

45) A compound contains 28% of Nitrogen and 72%of metal by weight. Three atoms of metal combineswith Two atoms of Nitrogen. The atomic weight of metal is

1) 12 2)24 3)36 4)48

46) An oxide of nitrogen contains 36.8% by weight of nitrogen. The formula of the compound is

1) N_2O 2) N_2O_3 3) NO 4) NO_2

47) A certain compound contains magnesium, carbon and Nitrogen in the mass ratio 12 : 12 : 14. The formula of the compound is

1) MgCN 2) Mg_2CN 3) $MgCN_2$ 4) $Mg(CN)_2$

48) At STP one litre of a gas weighs 1.25 grams. The gas contains 85.71 % of carbon and 14.29% of hydrogen. The formula of the compound is

1) CH_4 2) C_2H_6 3)C_3H_8 4) C_2H_4

KEY

1) 3	2) 1	3) 1	4) 4	5) 2	6) 2	7) 2	8) 3
9) 2	10) 4	11) 1	12) 2	13) 3	14) 1	15) 4	16) 2
17) 1	18) 1	19) 3	20) 4	21) 1	22) 4	23) 4	24) 2
25) 2	26) 4	27) 2	28) 3	29) 3	30 3	31) 3	32) 3
33) 4	34) 4	35) 2	36) 1	37) 2	38) 3	39) 3	40) 3
41) 4	42) 2	43) 4	44) 4	45) 2	46) 2	47) 4	48) 4

6. THERMODYNAMICS

Synopsis:-

1. **Thermodynamics**:- The subject dealing with Quantitative relation between heat energy and other forms of energy in physico-chemical processes.
2. There are three laws of Thermodynamics. These laws are based on experimental facts but not on the theoretical facts.
3. Thermodynamics predicts the energy transformations and feasibility of a process.
4. These laws do not give any idea about the rates of the processes.
5. The terms used in Thermodynamics are system and surrounding.
6. **System**:- It is any part of universe that is under thermodynamic study at that instant.
7. **Surroundings**:- The remaining part of the universe.
8. Systems are classified on the basis of their interaction with the surroundings as follows.
9. **Open System**:- The system where matter and energy are exchanged with surroundings. Eg. All living being, A cup containing water.
10. Closed System:- The system where only the energy but not the matter is exchanged with the surroundings. Eg. A perfectly insulated closed flask containing water.
11. Isothermal process:- The process that takes place at constant temperature.
    $\Delta T = O$
12. It is conducted very slowly in container made up of good thermal conductor.
13. Adiabatic Process:- The process in which there is no exchange of heat between system and the surroundings. (Δ Q=0)
14. Isobaric Process: - The process that takes place at constant volume (Δ V=0)
15. State function (or) state variable:- It is the property of a substance that depends on state of that substance but not on how that state is achieved. Eg. Internal energy, Enthalpy, Entropy, Gibb's energy, Temperature, Pressure, volume etc.
16. Path function:- The property of a substance that depends on the path i.e how that substance is derived. Eg. Work.
17. State of a system:- It is defined in terms of it's state functions such as P,V,T etc.
18. If any one of the state functions is changed, the state of that system is said to be changed.
19. Extensive Property:- It is the property of a substance that depends on it's mass. Eg. Volume of a gas, internal, energy, Enthalpy, entropy etc.
20. Intensive Property:- It is the property of a substance that does not depend on it's mass. Eg. density, melting point, boiling point, temperature, viscosity etc.
21. Cyclic process:- When a system under goes a number of different processes and finally reaches to it's initial state it is called a cyclic process.
22. In a cyclic process Δ E=0, Δ H=0, $\Delta G = 0$
23. Reversible process:- A process is said to be reversible if it is brought out in such a way that the process could at any moment, be reversed by infinite small change. A reversible process proceeds infinitely slowly by a series of equilibrium states such that the system and surroundings are always at near equilibrium with each other.

24. The work obtained is maximum in a reversible process.
25. Irreversible process:- When a process goes from initial state to final state in a single step in finite time and can't be reversed it is called the irreversible process. All the natural processes are spontaneous and irreversible.
26. Heat (Q):- It is the form of energy which flows between a system and surroundings by virtue of temperature difference.
27. When some heat brings only change in temperature. The Heat absorbed (or) released $\Delta Q = ms \Delta t$ m = mass of the substance; s = Specific heat;

    Δt = change of temp.
28. Work (w) :- Work is said to be done when an unbalanced force causes some displacement in it's own direction.
29. It is calculated as the product of external pressure and change of volume. $W = -(P \Delta V)$; ($\Delta V = V - V$);
30. 'W' is +ve when work is done on the system.
31. 'W' is –ve when work is done by the system
32. Work is a path function. The units of work are Joule, erg, cal etc.,
33. 1 Joul = 0.2390 cal; 1 cal = 4.18 J ; 1 lit-atm = 101.3J = 1.013×10^9 erg = 24.2 cal.
34. Internal energy (E (or) U) :- It is the sum of all types of potential and kinetic energies of constituent particles of a given substance.
35. It is an extensive property and a state function.
36. It is impossible to determine the absolute value of 'E' of a substance. But the change of internal energy of a system (ΔE) can be determined.

    ΔE = heat absorbed (or) released in a process at constant volume and temperature

    $\Delta E = Q_v$
37. ΔE of as chemical reaction is determined in a Bomb calorimeter.
38. First law of thermodynamics:- It was proposed by Robert Mayer of Helmholtz.It is another form of "law of conservation of energy". It can be stated as "The energy of an isolated system is constant whatever changes take place in it"(or) "It is impossible to construct a perpetual machine that can work without consuming any form of energy ". (or) "The net energy change in a closed system is equal to heat absorbed plus the work done by the system"
39. The mathematical form of first law is ($\Delta E = \Delta Q + W$) ΔE = Change of internal energy; Q = heat gained or lost by the system; W = Work done by the system (or) on the system.
40. For absorption of heat 'Q' is + ve and for release 'Q' is –ve.
41. When work is done on the system 'W' is +ve. When works is done by the system 'W' is –ve.
42. Enthalpy (H):- Heat content of a system at constant pressure is called it's enthalpy.
43. It is a state function if an extensive property. It is calculated as the sum of internal energy and the product of pressure of volume. $H = E + PV$
44. It is impossible to determine the obsolute enthalpy of a system. But the change of enthalpy ($H = H_{final} - H_{intial}$) of a process can be determined

ΔH = The heat of a process when it is conducted at constant pressure and constant temperature.

$\Delta H = Q_p$

ΔH of a process can be calculated as
$$\begin{aligned} \Delta H &= \Delta E + \Delta W \\ &= \Delta E + P \Delta V \\ &= \Delta E + \Delta nRT \end{aligned}$$

n = no of gaseous moles of products- no. of gaseous moles of reactants

45. Limitations of first law:-
    a) It does not predict the direction of a spontaneous process.
    b) It does not explain why most of the exothermic reactions do not proceed to completion.
    c) It does not explain why natural processes are irreversible.
46. Spontaneous process:- The reaction that occurs under given set of conditions.(or) “The process which has a natural tendency to occur either on its own or after proper initiation under a given set of conditions”.
47. A spontaneous process need not be instantaneous. Ex: Dissolution of sugar in water, Evaporation of water, Mixing of any two gases, Burning of coal, Melting of ice, Neutralisation reactions, Decomposition of $CaCO_3$
48. Entropy (S) :- “It is a Thermodynamic state function which is a measure of randomness (or) disorder of the particles of a system”. It is more convenient to use change of entropy (S) Mathematically $\Delta S = \Delta S_{final} - \Delta S_{initial}$

    $\Delta S = \Delta S_{products} - \Delta S_{reactants}$ (for a chemical reaction).
49. Entropy is a state function and an extensive property it is the unavailable energy in the system.
50. For a spontaneous process in an isolated system

    $\Delta S > 0$ i.e positive.
51. When a system is not isolated the entropy changes of the surroundings also must be considered. Then $\Delta S_{Total} = \Delta S_{system} + \Delta S_{surrounding}$
52. For a spontaneous process $\Delta S_{total} > 0$.
53. When ΔS is + ve the process is spontaneous, when ΔS is –ve the process is non-spontaneous, when $\Delta S = 0$ the process is in equilibrium.
54. The entropy of a system increases when it absorb heat.
55. ΔS is more +ve when the system absorbs heat at lower temp rather that at higher temp. Units of S and ΔS are $J.K^{-1}\ mol^{-1}$
56. At a given temp $S_{Gas} > S_{liquid} > S_{Solid}$.
57. Entropy of fusion:- It is the change in entropy when one mole of a solid changes to a liquid at it is melting point. Here heat absorbed (Q_{rev}) = Latent heat of fusion (H_{fusion})

    $\Delta S_{fusion} = S_{liquid} - S_{solid}$ = melting point (K)
58. Entropy of vapourisation:- It is the change in entropy when one mole of a liquid changes to vapour at it’s boiling point.

    $S_{vapour} = S_{vapour} - S_{liquid} = \dfrac{\Delta H_{vapour}}{\text{Boiling Point (K)}}$
59. Entropy of sublimation:- It is the change of entropy when one mole solid changes into vapour at a particular temperature.

    $\Delta S_{sub} = S_{vapour} - S_{solid} = \underline{\Delta H_{sub}}$

60. Second law of Thermodynamics:- It is stated in various forms. Some of them are as follows. Heat cannot flow from a colder body to a hotter body on its own. Heat cannot be converted in to work completely without causing some permanent changes in the system or in the surroundings.
61. All spontaneous processes are thermodynamically irreversible and the entropy increases.
62. It is impossible to construct a machine working in cycles and transfers heat from a lower temperature region to a higher temperature region without intervention of an external agency (such an imaginary machine is called perpetual motion machine of second kind)
63. Second law tells us whether a given process can occur spontaneously and to what extent. It also helps us to calculate the maximum fraction of heat that can be converted into work in a given process.
64. The mathematical form of second law is $\Delta S > \frac{\Delta Q}{T}$
65. For a reversible process $\Delta S = \Delta q / T$. For an irreversible process $\Delta S > \Delta q / T$
66. Gibb's energy (or) Gibb's function (G) :-"It is a thermodynamic Quantity of a system, the decrease of which is equal to useful work done by the system in a process"
67. Mathematically: $G = H - TS$ $\Delta G = \Delta H - T \Delta S$ (Gibb's – Helmholtz equation)
68. In a non – isolated system $\Delta G_{system} = T \Delta S_{total}$
69. It is the ultimate driving force of a spontaneous process which is indicated by –ve value of G. i.e. If $\Delta G < 0$ i.e -ve the process is spontaneous.
    If $\Delta G > 0$ i.e +ve the process is non spontaneous.
    If $\Delta G = 0$ i.e. the process is equilibrium state
70. Gibb's free energy and work: Gibb's free energy change of a process is equal to the maximum possible useful work that can be derived from that process.
    $\Delta G = W_{max}$ (for Reversible process at constant P&T)
71. In the case of Galvanic cells $\Delta G^0 = - nFE^0$ n = no. of electrons involved
    F = Faraday = 96,500c, E^0= std. emf of the cell.
72. For an equilibrium $G^0 = RT \ln k = -2.303\ RT \log_{10} k$, k= equilibrium constant.
73. Third law of thermodynamics:- It was proposed by Max plank W. Richard & Walter in different forms.
74. "The entropy of a pure and perfectly crystalline substance is zero at the absolute zero temperature. $S_{\lim T-0} = 0$
75. "It is impossible by any method no matter how idealized it is, to reduce the temperature of any system to the absolute zero in a finite number of operations"
76. Third law imposes a limitation on entropy value but not leads to any new thermodynamic concept.
77. Debye equation $C_v = aT^3$; a = Constant for a substance
78. At the vicinity of absolute zero $C_p = C_v$
79. Hence absolute value of 'S' can be calculated using C_v value.

Heat capacity and specific heat

- Heat capacity (c) of a substance is the amount of heat required to raise its temperature through one degree.
- Heat capacity is the heat absorbed by the system to the resulting increase in temperature. $C = \frac{q}{dt}$ q = heat absorbed

  Dt = rise in temperature
- Heat capacity is of two types for gases.

  a) Heat capacity at constant volume (C_V) give s the measure of change of internal energy (e) of a system with temperature. $C_V = \left[\frac{dE}{dT}\right]_V$

  b) Heat capacity at constant pressure (C_P) gives the measure of change of enthalpy (H) of a system with temperature. $C_P = \left[\frac{dH}{dT}\right]_P$

  $C_P = C_V + R \qquad \frac{C_P}{C_V} = r$

  Specific heat is the required to raise the temperature of one gram of a substance through $1^o C$

  Molar heat capacity (or) molar heat = specific heat x mol. Wt of substance

Measurement of ΔE *and* ΔH of Chemical Reaction

- The technique of measurement of heats of reactions is called calorimetry.
- The apparatus used is called calorimeter.
- Depending on types of chemical reactions, two types of calorimetry used.
- First type is used in combustion reactions.
- Second type is used in dissolution of solid in water and neutralization reactions.

First type (for combustion reactions)

- This type of calorimeter is knows as bomb calorimeter.
- The bomb is made of steel & coated inside with platinum or gold or some non-oxidisable material.
- A known weight of combustible substance is ignited by passing electric current through platinum wire. The substance undergoes combustion and heat liberated increases the temperature of water in the calorimeter. The rise in temperature is measured by Backmann thermometer.
- The heat capacity of calorimeter is determined using a known weight of benzoic acid prior to the main experiment.
- The heat of combustion of benzoic acid is 322 KJ / mole.
- Heat of combustion = $Z \times O \times \frac{M}{m} cals$

  O = rise in temperature of water in calorimeter after combustion.

  Z = heat capacity of calorimeter + water.

  m = Wt of substance in gms.

$q_p = q_v + \Delta nRT$ Δn = change in number of gas molecules.

q_v = heat of combustion at constant volume.

Second type of Calorimeter

- Dewar flask can be used as calorimeter.
- Water equivalent (W) of the calorimeter together with stirrer and the thermometer is measured.

$$W = \left[\frac{m_2(t_2 - t_3)}{t_3 - t_1} - m_1\right]$$

$t_2{}^oC$ = water at high temperature.

m_2 = known mass at temperature

$t_1{}^oC$ = water at lower temperature

m_1 =known mass at

$t_3{}^oC$ = mixed substance resultant temperature

W = mixed substance mass

- Heat liberated

  (W + Volume of reaction mixture) x rise in temperature.

Bond Dissociation Energy:

The amount of energy required to break mole of a particular bond in a given compound and to separate the resulting gaseous atoms or ions or radicals from one another is bond dissociation energy. $H_2 \to 2H \quad \Delta H^o + 435.9\,KJ/mole$

- The bond dissociation energy in polyatomic molecules will be only average value, because in each step of dissociation different fragment are involved.

$$CH_{4(g)} \to CH_{3(g)} + H_{(g)} \quad \Delta H^o = 427\,KJ$$

$$CH_{3(g)} \to CH_{2(g)} + H_{(g)} \quad \Delta H^o = 418.4\,KJ$$

$$CH_{2(g)} \to CH_{(g)} + H_{(g)} \quad \Delta H^o = 460.2\,KJ$$

$$CH_{(g)} \to C_{(g)} + H_{(g)} \quad \Delta H^o = 343.1\,KJ$$

C – H bond dissociation energy = $\frac{427 + 418 + 460 + 343}{4} = \frac{1648.7}{4} = 412J$

Enthalpy of Atomization: The heat required to dissociate one mole of simple molecule in the gaseous state into its constituent atoms is called enthalpy of atomization. This is an endothermic process.

$$H_2 \to 2H \quad \Delta H = 43.51\,KJ$$

$$N_2 \to 2N_{(g)} \quad \Delta H = 937.4\,Kg$$

Enthalpy of sublimation:

Sublimation: some substances in the solid state at room temperature are converted on heating into gaseous state.

$$Solid \xrightarrow{\Delta} gas$$

$$SolidT_2 \xrightarrow{\Delta} T_2\,vapour$$

This is an endothermic process.

The amount of heat required to convert on mole of a simple substance in the solid state into the gaseous state without decomposition of the substance.

Enthalpy of Phase Transition: The heat change involved in the change of phase (or) physical state of one mole of compound at atmosphere pressure is called enthalpy of phase transition.

$$C_S = C_{(g)} \quad \Delta H = 1439.2 KJ$$

$$S_{r\,\mathrm{hom}\,bic} \rightarrow S_{monoclinic} \quad \Delta H = +2.5 Kg$$

Enthalpy of ionization: The enthalpy change in the formation of an ion at unit activity from its elements in aqeous solution is enthalpy of ionization.

$$+_2 H_2 + aq \rightarrow H_{(aq)} + e^- \qquad \Delta H^o = 0.0 KJ$$

Question Bank:

1) Which of the following are true about a "system"

   1) Will do not have definite amount of substance

   2) Energy and matter may not be exchanged with surroundings

   3) (Universe+surroundings)

   4) (Universe-surroundings)

2) The object under thermodynamically study is called

   1) System 2) Universe 3) Surrounding 4) Boundary

3) Thermodynamic laws speak about

   1) Rates of chemical changes

   2) Feasibility and energy transformations of a process

   3) Both the rate and energy changes of a process

   4) Energy changes in chemical reactions only

4) In an open system

   1) Energy is exchanged with surroundings

   2) Matter is exchanged with surroundings

   3) Both energy and mater are exchanged with surroundings

   4) Neither energy nor matter are exchanged

5) "Closed system" is

   1) Perfectly sealed 2) Perfectly insulated

   3) Both perfectly insulated and sealed 4) Neither insulated nor sealed

6) In a closed system

   1) Energy is not exchanged 2) Matter is exchanged

   3) Energy is only exchanged 4) Energy and matter are exchanged

7) In open system, system and surroundings exchange

   1) Energy only 2) Matter only

   3) Energy and matte 4) Neither energy nor matter

8) Hot water in a thermos flask is an example for

   1) Isolated system 2) Open system

   3) Closed system 4) Adiabatic system

9) Which one of the following statement is false

   1) Work is a state function

   2) Temperature is a state function

   3) Change in the state is completely defined when the initial and final states are specified

   4) Work appears at the boundary of the system

10) Which of the following is a path function

1) Internal energy2) Enthalpy 3) Work 4) Entropy

11) According to 1st law of Thermodynamics

1) The energy of system is constant

2) The energy of universe is constant

3) The energy of surroundings is constant

4) The energy of system and surroundings are not constant

12) The limitation of 1st law of thermodynamics

1) Energy neither created nor destroyed

2) Impossible to construct the 1st kind of perpetual motion machine

3) Law of conservation of energy

4) Spontaneous process are irreversible

13) Which of the following statement is true about 1st law of Thermodynamics

1) Energy is created nor destroyed

2) If particular form of energy disappears the equivalent quantity appears in same from

3) It is possible to construct a perpetual motion machine of first kind

4) Energy neither created not destroyed

14) According to 1st raw of Thermodynamics

1) Energy can be created but not destroyed

2) Energy cannot be created but can be destroyed

3) Energy can be created and destroyed

4) Energy cannot be created nor destroyed

15) First law of Thermodynamics is also known as

1) Law of conservation of energy 2) Law of conservation of mass

3) Law of multiple proportions 4) Law of dis proportionation

16) The work done when a gas is compressed by an average pressure of 0.50 atm so as to decrease its volume from 400cm^3 to 200cm^3

1) 10.13 J 2)20.13 J 3) 30.13 J 4) 40.13 J

17) A gas expands by 0:5 litre against a constant pressure of one atmosphere. Then the work done in Joules

1) -50.65 J 2) -100.5 J 3) -200 J 4) -10 J

18) 5L cylinder contains 10 mol of nitrogen at 27°C If whole of the gas escapes into atmosphere having pressure 1atm. then the work has to be done by the gas is

1) -10K.J 2) -15 K.J 3) -24.413 K.J 4) -50.4 K.J

19) A gas expands by 0.5L against a constant pressure of 1 atm. then the work done in calories is

1) -50.664 cal 2) -12.10 cal 3) -5.64 cal 4) -35.61 cal

20) When a sample of gas expands from 4.0L to12.0 L against a constant pressure of 0.30 atm, the work involved is

1) 243.19 J 2) -243.19 J 3) 234.19 J 4)-234.19 J

21) A gas expands from 1.5 to 6.5 L against a constant pressure of 0.5 atm and during this process the gas also absorbs 100 J of heat. The change in the internal energy of the gas is

1) 153.3 J 2) 353.3 J 3) -153.3 J 4) -353.3 J

22) If a gas expands adiabatically from 1.0 L to 12.0 L against a constant pressure of 0.75 atm. The ΔU of the gas is

1) -853.9 J 2) 835.9 J 3) 853.9 J 4) -835.9 J

23) The difference between heat of reaction at constant pressure and constant volume for the reaction given below at 25°C in kJ is:

$2C_6H_{6(l)} + 15O_{2(g)} \rightarrow 12CO_{2(g)} + 6H_2O_{(l)}$

1) -7.43 2) +3.72 3) -3.72 4) +7.43

24) A system absorbs 20 KJ of heat and also doc 1 OKJ of work. The net internal energy of the system

1) Increases by 10 KJ 2) Decreases by 10 KJ

3) Increases by 30 KJ 4) Decreases by 30 KJ

25) If a gas absorbs 200 J of heat and expands by 500 cm^3 against a constant pressure of 2×10^5 Nm^{-2}, then change in internal energy is

1) -300 J 2) -100 J 3) +100 J 4) +300 J

26) A system performs 101.3 J of work on its surrounding arid absorbs 15 KJ of heat from surroundings. The change in internal energy of the system

1) 15101.3 J 2) 14898.7 J 3) 13972.5 J 4) 12496.3 J

27) One mole of liquid water at is boiling point vaporises against a constant external pressure of 1 atm.at the same temperature. Assuming ideal behaviour and initial volume of water vapours are zero, the work done by the system nearly

1) -3102 J 2) +3102 J 3) -4268 J 4) + 4268 J

28) A system absorbs 10 kJ of heat and does 4 kJ of work. The internal energy of the system

1) Decreased by 6 kJ 2) Increases by 6 kJ

3) Decreases by 14 kJ 4) Decreases by 14 kJ

29) Which of the following is an endothermic reaction

1) $N_2(g) + 3H_2(g) - 92\ kJ \rightarrow 2NH_3(g)$

2) $N_2(g) + O_2(g) + 180.8\ kJ \rightarrow 2NO(g)$

3) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g) + 184.6\ kJ$

4) $C(gr) + 2H_2(g) \rightarrow CH_4(g) + 74.8\ kJ$

30) When 10 grams of methane is completely burnt in oxygen, the heat evolved is 560 kJ. What is the heat of combustion (in kJ.mol^{-1}) of methane?

1) -1120 2) -968 3) -896 4) -560

31) Which one of the following is an exothermic reaction?

1) $N_2(g) + O_2(g) + 180.8$ kJ $\rightarrow 2NO(g)$

2) $N_2(g) + 3H_2(g) - 92$ kJ $\rightarrow 2NH_3(g)$

3) $C(gr) + H_2O\ (g) \rightarrow CO(g) + H_2(g) - 131.4$ kJ

4) $C(gr) + 2S(s) \rightarrow CS_2(l) - 91.9$ kJ

32) Which of the following is an example of exothermic reaction?

1) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$; $\Delta H = -184.6$ KJ

2) $N_2(g) + O_2(g) \rightarrow 2NO(g)$; $\Delta H + 180.8$ KJ.

3) C (graphite) + H_2O (g) $\rightarrow$ CO (g) + H_2 (g) -1314 KJ.

4) C (graphite) + 2 S (s) + 91.9 KJ $\rightarrow CS_2(l)$

33) The difference in ΔH and ΔE for the combustion of methane at 27°C would be

1) -1800 Cal 2) -162 Cal 3) -1200 Cal 4) 0

34) For the reaction $C_2H_4\ (g) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 2H_2O(l)$; DE = -1415 KJ. The ΔH at 27°C is

1) -1410 KJ 2) -1420 KJ 3) +1420 KJ 4) +1410 KJ

35) One mole of ideal gas expands freely at 310 K from five litre volume to 10 litre volume. Then Δ E and Δ H of the process are respectively

1) 0 and 5 cal 2) 0 and 5 x 300 cal

3) 0 and 0 4) 5 and 0 cal

36) The standard heat of formation values of $SF_6(g)$,S(g) and F(g) are -1100, 275 and 80 J mol^{-1} respectively. The energy S-F bond energy in SF_6

1) 309.2KJ 2) 378.2KJ 3) 284.6 KJ 4) 261.8 KJ

37) The heat of dissociation (in K.cals/mole) of CH_4 and C_2H_6 are 360 and 620 respectively. From these the C - C bond energy in the ethane can be evaluated as

1) 260 2) 130 3) 80 4) 200

38) The heat of formation (ΔH_f) of $H_2O(g)$ is -243 KJ. Δ E of it at temperature T is

1) -243 + RT 2) -243 – RT 3) -243 + 1/2 RT 4) -243 - 1/2 RT

39) Which of the following is not correct

1) Dissolution of NH_4Cl in excess of water is an endothermic process

2) Neutralization process is always exothermic

3) The absolute value of enthalpy (H) can be determined experimentally.

4) The heat of reaction at constant volume is denoted by D E

40) The relationship between ΔH and ΔE for the reaction PCl_3 (g) + Cl_2 (g) → PCl_5 (g) is given as

1) ΔH = ΔE + RT 2) ΔH = ΔE - RT

3) ΔH = ΔE - 2RT 4) ΔH = ΔE + 2RT

41) The heat of combustion of methane at 298 K is expressed by

CH_4 (g) + $2O_2$ (g) → CO_2 (g) + $2H_2O$(l); ΔH = -890.2 KJ ΔE of the reaction at this temperature is

1) Less than ΔH 2) More than ΔH

3) Equal to ΔH 4) Not equal to ΔH

42) In the equilibrium reaction $N_2 + 3H_2 \rightarrow 2\ NH_3$ the sign of ΔH accompanying the reaction is;

1) Positive 2) Negative 3) Positive or Negative 4) Zero

43) If ΔE is the heat of reaction for $C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$ at constant volume, the ΔH (Heat of reaction at constant pressure) at constant temperature is

1) ΔH = ΔE + 2RT 2) ΔH = ΔE - 2RT

3) ΔH = ΔE + RT 4) ΔH = ΔE - RT

44) For the reaction, $C + O_2 \rightarrow CO_2$

1) ΔH > ΔE 2) ΔH < ΔE 3) ΔH = ΔE 4) ΔH = 0

45) For the system S (s) + O_2 (g) → SO_2 (g)?

1) ΔH > ΔE 2) ΔE > ΔH 3) ΔH = 0 4) ΔH = ΔE

46) Equal volumes of methanoic acid and sodium hydroxide are mixed. If x is the heat of formation of water, the heat evolved on neutralisation is

1) More than x 2) Equal to x 3) Twice of x 4) Less than x

47) If x K.cal and y Kcals are the ΔH values of ionisation of weak acid (WA) and weak base (WB) respectively, ΔH of neutralisation of the WA and WB is

1) 13.7 K.cal 2) -13.7 + (x-y) K.cal

3) -13.7 + (x + y) K.cal 4) -13.7 - (x + y) K.cal

48) The heat of neutralisation of NaOH with HCI is 57.3 KJ and with HCN is 12.1 KJ. The heat of ionisation of HCN is

1) +69.4 KJ 2) +45.2 KJ 3) -45.2 KJ 4) -69.4 KJ

49) Equal volumes of molar HCI and H_2SO_4 are neutralised by dilute NaOH solution and x K.cal and y K.cal of heat are liberated respecitvely .Which of the following is true?

1) x = Y 2) x = y/2 3) x = y/3 4) x = 2y

50) If 9.8 gms, of H_2SO_4 is neutralised exactly with strong alkali, the heat evolved is (ΔH of neutralisation is 13.6 K.cal)

1) 13.6 K.cal   2) 1.36 K.cal   3) 2.72 K.cal   4) 27.2 K.cal

51) If $H^+ + OH^- \rightarrow H_2O.\Delta H = -57.3KJ$, then the heat of neutralisation for complete neutralisation of 1 mole H_2SO_4 by a base is

1) -57.3 KJ   2) -114.6 KJ   3) -28.65 KJ   4) -14.3 KJ

52) The standard heat of formation of carbon disulphide (*I*) given that standard heat of combustion of carbon (s), sulphur (s) and carbon disulphide (*I*) are -393.3, -293.72 and - 1108.76 KJ mol^{-1} respectively is

1) -12.502 KJ. mol^{-1} 2) +128.02 $KJ.mol^{-1}$

3) -128.02 $KJ.mol^{-1}$ 4) +12.802 $KJ.mol^{-1}$

53) The heat of formation of the compound in the following reaction is

$H_2 + CI_2 \rightarrow$ 2 HCI + 44 K.Cal

1) -44 K.cal   2) -22 K.caf   3) +11 Kcal   4) -88 K.cal

54) In order to decompose 9 grams of water 142.5 KJ heat is required. Hence the enthalpy of formation of water is

1) +142.5 KJ   2) -142.5 KJ   3) +285 KJ   4) -285 KJ

55) The heat of formation of $CO_{(g)}$ and $CO_{2(g)}$ are ΔH= 110 and ΔH= -393 $kJ.mol^{-1}$ respectively. What is the heat of reaction (ΔH) (in $kJ.mol^{-1}$) for the following reaction

1) -504   2) -142.5   3) -283   4) 504

56) Energy required to dissociate 4 gms of H_2 (g) into free atoms is 208 K.cals at 25°C. The H-H bond energy will be

1) 104 K.cals   2) 1040 K.cals   3) 10.4 K.cals   4) 208 K.cals

57) The enthalpy change when 1.00 g of water freezes at 0°C and 1.00 atm is (ΔH_{fusion} = 1.435 K.cal/mol)

1) 0.079 cal/g   2) -79.2 cal/g   3) 99.7.7 cal/g 4) -79.7 cal/g

58) Given that $C_{(graphite)} \rightarrow C_{(g)}$; DH = +716.7KJ. $C_{(diamond)} \rightarrow C_{(g)}$; DH - 714.8KJ. The ΔH for the following reaction is $C_{(graphite)}$ $C_{(diamond)}$

1) 1.9 KJ   2) -1.9 KJ   3) Zero   4) 714.8 KJ

59) At 25°C the heat of formation of $H_2O(I)$ is -285.9 KJ $mole^{-1}$ and that for H_2O (g) is -242.8 $KJmole^{-1}$ The heat of vaporization of water at the same temperature is

1) 43.1 KJ $rnole^{-1}$ 2) 242.8 KJ $mole^{-1}$

3) -43.1 KJ $mole^{-1}$ 4)-242.8 KJ $mole^{-1}$

60) Δ x is given by: Δ x = (hydration energy) - (lattice energy). For a salt, Δ x is found to be 500 k cal mol^{-1}. Then on increasing the temperature, Its solubility in water

1) increases 2) decreases
3) remains same 4) increases initially and then decreases

61) Δ x is given by: Δ x = (hydration energy) - (lattice energy). For a salt, D x is found to be 500 k cal mol^{-1}. Then on increasing the temperature, its solubility in water

1) increases 2) decreases
3) remains same 4) increases initially and then decreases

62) The enthalpy of KNO_3 is 36.5 K.cals. In this reaction

1) Energy is liberated 2) Energy is absorbed
3) KNO_3 is stable4) KNO_3 is unstable

63) Given that $Zn + \frac{1}{2}O_2 \rightarrow ZnO + 35.25KJ$.

$HgO \rightarrow Hg + \frac{1}{2}O_2$ -9.11KJ. The heat of the reaction $Zn + HgO \rightarrow ZnO + Hg$ is

1) -26.14 KJ 2) 44.39 KJ 3) -44.39 KJ 4) 26.14 KJ

64) $H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$, DH = -44 K.cals

$2\,Na(s) + 2\,HCl(g) \rightarrow 2\,NaCl(s) + H_2(g)$, $\Delta H = -152$ K.cal

$Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$, $\Delta H = ?$

1) +108 K.cal 2) -196 K.cal 3) -98 K.cal 4) 50 K.cal

65) If $S(g) + e^- \rightarrow S^{1-}(g)$, $\Delta H = -207.6$ KJ

$S(g) + 2e^- \rightarrow S^{2-}(g)$, $\Delta H = +335.2$KJ. The enthalpy for the reaction

$S^{1-}(g) + e^- \rightarrow S^{2-}(g)$:

1) +127.6 KJ 2) 127.6 KJ 3) +542.8 KJ 4) 641.8 KJ

66) When enthalpy and entropy change for a chemical reaction are -2.5 x 10^3 cals and 7.4 cals deg^{-1} respectively. Predict that reaction at 298 K is

1) Spontaneous 2) Reversible 3) Irreversible 4) Non-spontaneous

67) If ΔH_{vap} of pure water at 100°C is 40.627 kJ mol^{-1} The value of ΔS_{vap} is

1) 108.91 kJmol^{-1} 2) 108.91 JK^{-1} mol^{-1}
3) 606.27 JK^{-1}mol^{-1} 4) 808.27 JK^{-1}mol^{-1}

68) The temperature at which the reaction $Ag_2O(s) \rightarrow 2Ag(s) + \frac{1}{2}O_2(g)$ at 1atmospheric pressure will be in equilibrium is __ K. The value of ΔH and ΔS for the reaction are 0.58 KJ and 66.11 JK^{-1} respectively and these calue do not change much with temperature.

1) 462.6 K 2) 486.4 K 3) 364.5 K 4) 521.2 K

69) The latent heat of fusion of ice is 5.99 KJ/mol at its melting point. Then

(i) AS for fusion of 900 g ice and

(ii) AS for freezing of liquid water are respectively

1) 8698 JK^{-1} ; -16.8 J $mol^{-1}K^{-1}$ 2) 9269 J K^{-1} ; -12.6 J mol^{-1} K^{-1}

3) 1097 JK^{-1} ; -21.9 J $mol^{-1}K^{-1}$ 4) 1236JK^{-1} ; -28.6 J mol^{-1} K^{-1}

70) Combustion of hydrogen in a fuel cell at 300 K J is represented as $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$. If Δ H and Δ G are-241.60 KJ mol^{-1} and -228.40 KJ mol^{-1} of H_2O. The value of Δ S for the above process is

1) 4.4 kJ 2) -88 j 3) +88 J 4) -44 J

71) In a chemical reaction, ΔH = 150K.J and ΔS = 100JK^{-1} at 300K. ΔG would be

1) Zero 2) 300KJ 3) 330KJ 4) 120KJ

72) The change in entropy, ΔS is positive for an endothermic reacition, If enthalpy change ΔH occurs at the same temperature T, then the reaction is feasible

1) At all temperatures 2) When $\Delta H > T\ \Delta S$

3) When $\Delta H < T\ \Delta S$ 4) Not feasible at all

73) Standard enthalpy and standard entropy for the oxidation of NH_3 at 298K are -382.64K.Jmol^{-1} and -145.6 J.mol^{-1} respectively. Standard free energy change for the same reaction at 298K is

1) -221.1KJmol^{-1} 2) -339.3 K.Jmol^{-1}

3) -439.3 K.Jmol^{-1} 4) -523.2 K.Jmol^{-1}

74) The ΔG in the process of melting' of ice at -15^{0}C is

1) ΔG is –Ve 2) ΔG is +Ve 3) $\Delta G = 0$ 4) All of these

75) For a reaction $A_{(g)} \rightleftharpoons B_{(g)}$ at equilibrium the partial pressure of B is found to be one fourth of the partial pressure of A. The value of ΔG^0 of the reaction A → B is

1) RT ln 4 3) -RT In 4 2) RT Log 4 4) -RT log 4

76) For a reaction 2 A (s) + 2 B (g) → 2 C(I) + D(g). If heat of reaction at constant pressure is -28 K.cal. The heat of reaction at constant volume at 27^{0}C is

1) -27.4 K.cal 2) 27.4 K.cal 3) 28.4 K.cal 4) 28 K.cal

77) When 10 grams of methane is completely burnt in oxygen, the heat evolved is 560 kJ. What is the heat of combustion (in kJ.mol^{-1}) of methane?

1) -1120 2) -968 3) -896 4) -560

78) The enthalpies of formation of CO and CO_2 are -110.5 KJ mol^{-1} and -393.5 KJ mol^{-1} respectively. The enthalpy of combustion of carbon monoxide is

1) -110.5 KJ mol^{-1} 2) -283.0 KJ mol^{-1} 3) 504.0 KJ mol^{-1} 4) 221.2 KJ mol^{-1}

79) The heats of combustion of C_2H_4 , C_2H_6 and H_2 gases are -1409.5 KJ, -1558.3 KJ and -285.6 KJ respectively. The heat of hydrogenation of ethene is

1) -136.8 KJ 2) -13.68 KJ 3) 273.6 KJ 4) 1.368 KJ

80) When 6 g carbon is burnt in a sufficient amount of oxygen, the heat evolved is x KJ. The heat of combustion of carbon is

1) x KJ 2) 2x KJ 3) 4x KJ 4) None

81) The heats of combustion for C, H_2 and CH_4 are -349, -241.8 and -906.7 KJ respectively. The heat of formation of CH_4 is

1) 174.1 KJ 2) 274.1 KJ 3) 374.1 KJ 4) 74.1 KJ

82) The DH for the conversion of C (diamond) to C (graphite) when the following reactions are given C (diamond) + $O_2(g) \rightarrow CO_2(g)$; $\Delta H = -94.5$ Kcal.

C (graphite) + $O_2(g) \rightarrow CO_2(g)$; DH = -94.0 K.cal.

1) -188.5 K.cal 2) +188.5 K.cal 3) +0.5 Kcal 4) -0.5 K.cal

83) $H_2(g) + I_2(s) \rightarrow 2HI(g); \Delta H = 51.9$ KJ/mole

$H_2(g) + I_2(g) \rightarrow 2HI(g); \Delta H = -9.2$KJ/mole. The heat of reaction of $I_2(s) \rightarrow I_2(g)$

1) -61.1 KJ 2) +61.1 KJ 3) +31.1 KJ 4) 50.1 KJ

84) AT 298 K, the heat of solution of $MgSO_4$ (s) is -91.21 KJ mol^{-1} and that of $MgSO_4.7H_2O$ (s) is 13.81 KJ mol^{-1}. Calculate heat of hydration of $MgSO_4$ (s) i.e., ΔH for the reaction

$MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4.7H_2O(s)$

1) -105.02 KJ mol^{-1} 2) 105.02 KJ mol^{-1}

3) -77.40 KJ mol^{-1} 4) 77.40 KJ mol^{-1}

85) The value 11 S for the process $H_2O(s) \rightarrow H_2O(I)$ at 1 atm pressure and 260K is greater than zero. The value of Δ G will be

1) > 0 2) < 0 3) = 0 4) lies between-1 and 0

KEY

1) 4	2) 1	3) 2	4) 3	5) 1	6) 3	7) 3	8) 1
09) 1	10) 3	11) 2	12) 4	13) 4	14) 4	15) 1	16) 1
17) 1	18) 3	19) 2	20) 2	21) 3	22) 4	23) 1	24) 1
25) 3	26) 2	27) 1	28) 2	29) 2	30) 3	31) 2	32) 1
33) 3	34) 2	35) 3	36) 1	37) 3	38) 3	39) 3	40) 2
41) 2	42) 2	43) 4	44) 3	45) 4	46) 4	47) 3	48) 2
49) 2	50) 3	51) 2	52) 2	53) 2	54) 4	55) 3	56) 1
57) 4	58) 1	59) 1	60) 2	61) 2	62) 2	63) 1	64) 3
65) 3	66) 1	67) 2	68) 1	69) 3	70) 2	71) 4	72) 3
73) 2	74) 2	75) 1	76) 1	77) 3	78) 2	79) 1	80) 2
81) 4	82) 4	83) 2	84) 1	85) 1			

7a. CHEMICAL EQUILIBRIM

1. Irreversible reactions: In these reactions products cannot be converted into reactants.

   Ex: $2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$
   $NH_4NO_{2(s)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$
   $C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$
   $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$
   $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
   $H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$

2. Reversible reactions :
   - A reaction is said to be reversible, if the reaction mixture contains the reactants as well as the products and if both the forward and the backward reactions are taking place simultaneously under the given experimental conditions.
   - Reactants giving rise to products is known as forward reaction.
   - Products giving rise to reactants is known as forward reaction.
   - Reversible reactions are represented by writing a pair of half headed –arrows pointing in opposite directions in between the reactants and products.
   - A reversible reaction does not go to completion.
   - Most of the reversible reactions are carried in the closed vessels.

   Ex: $H_{2(g)} + I_{2(g)} \Leftrightarrow 2HI_{(g)}$
   $PCl_{5(g)} \Leftrightarrow PCl_{3(g)} + Cl_{2(g)}$
   $2NO_{2(g)} \Leftrightarrow N_2O_{4(g)}$
   $N_{2(g)} + O_{2(g)} \Leftrightarrow 2NO_{(g)}$
   $2SO_{2(g)} + O_{2(g)} \Leftrightarrow 2SO_{3(g)}$
   $CaCO_{3(s)} \Leftrightarrow CaO_{(s)} + CO_{2(g)}$
   $CH_3COOH_{(l)} + C_2H_5OH_{(l)} \Leftrightarrow CH_3COOC_2H_{5(l)} + H_2O_{(l)}$

   - A reversible decomposition reaction is called dissociation reaction.
   - Some reversible reactions behave as irreversible reactions and go to completion under some specific conditions.
   - The removal of one of the products of a reversible reaction leads to the completion of the forward reaction.

     Chemical equilibrium is called a dynamic equilibrium because at equilibrium, the reactions do not stop. The forward and backward reactions proceed with equal speed.

3. Till equilibrium is reached, there is a change in the detectable or macroscopic properties of the system. Once equilibrium is established, there is no further change in the macroscopic properties of the system at constant temperature and pressure.
4. Law of mass action states that the rate of a chemical reaction is proportional to the product of the active masses of the reacting substances.
5. Active mass is taken as equal to the molar concentration expressed in moles/litre in the case of gases or substances in solution.
6. The active mass of a solid is unity irrespective of its mass.

7. When all the reactants and products are in the same phase, the equilibrium is called a homogeneous equilibrium.

   Ex: Homogeneous gaseous equilibria

   $H_{2(g)} + I_{2(g)} \Leftrightarrow 2HI_{(g)}$

   Homogeneous equilibrium in solution

   $CH_3COOH_{(l)} + C_2H_5OH_{(l)} \Leftrightarrow CH_3COOC_2H_{5(l)}+H_2O_{(l)}$

8. When the reactants and products are in different phases, the equilibrium is called a heterogeneous equilibrium.

   Ex: $CaCO_{3(s)} \Leftrightarrow CaO_{(s)}+CO_{2(g)}$

9. Equilibrium can be approached from either direction i.e, either by starting with the reactants or products.
10. Once equilibrium is reached, it continues forever until the conditions like pressure, temperature etc, are altered.
11. Equilibrium constant $K = \frac{K_f}{K_b}$

    Where K_f and K_b are the velocity constants of the forward and backward reactions respectively.

12. K_f and K_b are rate constants. They depend upon temperature and the nature of the reaction. Hence the equilibrium constant for a given reaction and is independent of pressure, concentration, volume and the presence or absence of a catalyst or a foreign substance.
13. The rate constant of a reaction increases in presence of a catalyst but the equilibrium constant is independent of the presence or absence of catalyst.
14. In the case of gaseous equilibria, since the partial pressure of a gas is proportional to the molar concentration, molar concentrations can be replaced by partial pressures. Then for the reaction,

    $$QA+mB \Leftrightarrow nC+pD$$

    $$Kp = \frac{P_C^n \,.\, P_D^p}{P_A^Q . P_B^m}$$

15. For a homogeneous gaseous equilibrium $K_p = K_c(RT)^{\Delta n}$

    Where Δ = number of moles of gaseous products-number of moles of gaseous reactants.

    If Δn is positive, K_p is greater than K_C

    If Δn is negative, K_p is less than K_C

    If $\Delta n = 0$, $K_p= K_C$

16. In any spontaneous chemical reaction there is a decrease in free energy.

    i.e, ΔG is negative.

17. At equilibrium there is no change in free energy i.e, $\Delta F = 0$ or $\Delta G =0$
18. When equilibrium is attained at a high temperature, if it is suddenly cooled to laboratory temperature both forward and backward reactions almost stop and the equilibrium concentrations are arrested. This is known as freezing the equilibrium or arresting the equilibrium.
19. The numerical value of the equilibrium constant indicates the extent of the forward reaction by the time equilibrium is reached. If the equilibrium constant is

large, it indicates that the forward reaction has taken place to a greater extent before the equilibrium is reached.

20. In the case of an exothermic reaction, with an increase in temperature, the equilibrium constant decreases.
21. In the case of an endothermic reaction, with an increase in temperature, the equilibrium constant increases.
22. Le-Chatelier's principle states that "When the conditions of equilibrium like pressure, temperature and concentration are altered, the equilibrium moves in such a direction as to cancel or tend to cancel the effect of the change".
23. According to Le-chatlier, when temperature is increased, the position of equilibrium moves in the endothermic direction.
24. When pressure is increased, the equilibrium moves in a decrease in number of gaseous moles or decrease in volume.
25. When the concentration of the reactants is increased, the forward reaction is favoured.
26. The melting point of ice decreases with an increase in pressure because ice melts with a decrease in volume.
27. Synthesis of NH_3 by Haber's process:
    - Formation of ammonia is accompanied by decrease in volume and is favoured by high pressure.

      $N_{2(g)} + 3H_{2(g)} \Leftrightarrow 2NH_{3(g)} + 92$ Kj
    - The optimum pressure required is 200 atm.
    - The forward reaction is exothermic and favoured by low temperatures.
    - The low temperature will effect the rate of reaction hence an optimum temperature of 725-775K is maintained. (500^0C)
    - Finely divided Fe is used as a catalyst with Mo or the oxides of K & Al as a promoter.
28. Synthesis of Sulphuric acid by Contact process.
    - The forward reaction is accompanied by decrease in volume hence high pressure favours the forward reaction.

      $$2SO_{2(g)} + O_{2(g)} \Leftrightarrow 2SO_{3(g)} + 189.5 \text{ Kj}$$
    - An optimum pressure of 1 – 1.5 atm is used since at high pressure the acid corrodes the towers.
    - Low temperature favours the forward reaction since the forward reaction is exothermic.
    - The optimum temperature of 673 K is used. (400^0C)
    - Platinised asbestos or V_2O_5 is used as a catalyst.

QUESTION BANK

1) The effect of increasing the pressure on the following gaseous equilibrium $2A + 3B \Leftrightarrow 3C + 2D$ is

1) favours forward reaction 2) favours backward reaction

3) no effect 4) Favours for forward and backward reaction

2) At constant pressure, the additio o a gon

1) Reduces the formation of ammonia form nitrogen and hydrogen

2) Increases the formation of ammonia from nitrogen and hydrogen

3) Does not affect the equilibrium of the reaction in which ammonia is formed from nitrogen and hydrogen

4) Reduces the dissociation of ammonia

3) For N (g) + 3H (g) $\Leftrightarrow$ 2NH (g) + heat

1) K = K 2) K = K x (RT)

3) K = K x (RT) 4) K_p = K x $(RT)^{-1}$

4 For the equilibrium reaction 3Fe(s) + $4H_2O(g)$ $\Leftrightarrow$ Fe 0/s) +$4H_2(g)$ the relation between K_p & K_c is

1) K_p > K 2) K_p < K 3) K = K x $(RT)^{-}$ 4) K = K

5) The value of (K / K) for the reversible reaction SO (g) + 1/2 $O_2(g)$ $\Leftrightarrow$ SO (g) at constant temperature T is

1) (RT) 2) RT 3) (RT) 4) 1/RT

6) A reversible reaction A $\Leftrightarrow$ C takes place in two steps A $\Leftrightarrow$ B B $\Leftrightarrow$ C, which are also equilibrium steps If the equilibrium constants of the two steps are K_1 & K respectively the overall equilibrium constant K is equal to

1)K_1/K 2)K /K_1 3)K_1x K_2 4) K_1 - K_2

7) For the reaction 3O $\Leftrightarrow$ 2O

1) Kp > Kc 2) Kp < Kc 3) Kp = Kc 4) can not be predicted

8) For the homogeneous reaction 4NH (g) + 50(g) $\Leftrightarrow$ 4NO(g) + 6H O(g) t e equilibrium constant K_c has the unit of

1) $(Conc)^{-1}$ 2) Conc 3) $(Conc)^{+10}$ 4) it is dimensionless

9) For the reaction $H_2(g) + I_2(g)$ $\Leftrightarrow$ 2HI(g) at 721K the value of equilibrium constant is 50 The value of K_p under the same condition will be

1) 0.02 2) 0.2 3) 50 4) 50/RT

10) In the reaction H (g) + 1 (g) $\Leftrightarrow$ 2HI(g)

1) K = K 2) K ≠ 3) K_p > K 4) K < K

Fo e ea o g) + I (g) $\Leftrightarrow$ 2HI(g) at 721 K the li ium constant Kc is 50. When the li u oncentration of both is 0 5 M value o er he same conditions will be

0 002 3) 50 0 4) 50/RT

12) Forthe eac o PC (g) + Cl_2(g) $\Leftrightarrow$ PCls(g) The value o a 50 C is 26 The value of KR at the same e pe atu e will be

1) 0 0006 .57 3) 0 61 4) 0 83

13) At 27 C he value for the reaction PCl $\Leftrightarrow$ C 1 is 0 65 Its K is

1 .65 3) 0.2 4) 0 026

14) For the eactio (g) $\Leftrightarrow$ 8(g) + C(g), Kp at 400° C is .5 x 0 a at 600 C is 6 x 10^{-3} Then the incor ec s a ement is

1) the eaction is an exothermic one

2) Increase in temperature increases the formation of B

3) lnc ease i p essure increase the formation of A

4) A decrease in emperature and increase in pressure shifts the equilibrium towards left

15) for the synthesis of HI is 5 K for the dissociation of HI

1) 50 2) 5 3) 0.2 4) 0 02

16 In a chemica equilibrium the rate constant of the backward react on is 7 5 x 10^{-} and the equilibrium constant is 1 5. So the rate constant of the forward reaction is

1) 5 x 10^{-} 2) 2 x 10^{-3} 3) 1 125 x 10 4) 9 x 10^{-7}

17) An equilibrium mixture for the reaction 2H S(g) $\Leftrightarrow$ 2H (g) + S (g) had 1 mole of H S 0 2 mole of H and 0.8 mole of S_2 in 2 lit. flask. The value of Kc in mol.1it^{-1} is

1) 0 004 2) 0 08 3) 0.016 4) 0.16

18) In the gas phase reaction C H + H $\Leftrightarrow$ C H the equilibrium constant can be expressed in units of

1) Lit^{-} mole^{-1} 2) Lit. Mole 3) Mole Lit.$^{-2}$ 4) Mole Lit^{-1}

19) The equilibrium constant for the reaction N + 3H $\Leftrightarrow$ 2NH is K, then the equilibrium constant for the equilibrium NH $\Leftrightarrow \frac{1}{2}N_2+\frac{3}{2}H_2$

$\sqrt{K}$ $\frac{1}{\sqrt{K}}$

If K_1 and K_2 are the equilibrium constants for the reaction (1) and (2) respectively

N + O $\Leftrightarrow$ 2NO ...... (1);

NO $\Leftrightarrow \frac{1}{2}N_2+\frac{1}{2}$

1) $K_1=\left(\frac{1}{K_2}\right)^2$ 2) $K_1 = K_2^2$ 3) $K_1=\frac{1}{K_2}$ 4) $K_2=\frac{1}{K_1^2}$

21) At a given temperature the equilibrium constant for the reaction
$PCl_2(g) \Leftrightarrow PCl\ (g) + Cl\ (g)$ is 2.4×10^{-3}. At the same temperature the equilibrium constant for the reaction $PCl\ (g) + Cl\ (g) \Leftrightarrow PCl_2(g)$

1) 2.4×10^{-3} 2) -2.4×10^{-3} 3) 4.2×10^{-2} 4) $4\ 8\times10^{-}$

22) For the equilibrium the value of equilibrium constant K is $3.75\ 110^{-6}$ at 1069K. Calculate reaction at this temperature

1) 0 066 2) 3.33 3) 0.33 4) 0 033

23) Active mass of 56g of N_2 contained in 2 ltr. flask is

1) 16 Mole. lit^{-1} 2) 32 Mole. lit^{-1} 3) 1 00 mole lit^{-1} 4) 0 1 Mole lit^{-1}

24) What is the equation for the equ librium constant (K_c) for the following reaction

$$\frac{1}{2}A(g)+\frac{1}{3}B(g)\Leftrightarrow\frac{2}{3}C(g)$$

1) $K_c=\frac{[A]^{1/2}[B]^{1/3}}{[C]^{3/2}}$ 2) $K_c=\frac{[C]^{3/2}}{[A]^2[B]^3}$

3) $K_c=\frac{[C]^{2/3}}{[A]^{1/2}[B]^{1/3}}$ 4) $K_c=\frac{[C]^{2/3}}{[A]^{1/2}+[B]^{1/3}}$

25) In which of the following reactions, the concentration of product is higher than the concentration of reactant at equilibrium? (K = equilibrium constant)

1) $A \Leftrightarrow B$; K = 0.001 2) $M \Leftrightarrow N$ K = 10

3) $\Leftrightarrow Y$ K = 0 005 4) $R \Leftrightarrow P$; K = 0 01

26) I n t e reaction $PCl_s\ (g) \Leftrightarrow PCl\ (g) + Cl\ (g)$ the equilibrium concentration of PCl and PCl are 0.4 and O_2 molellitre respectively. If the value o is 0.5 what is concentration of Cl in moles/litre?

) 2 0 2) 1.5 3) 1.0 4) 0.5

27) E uilibrium constant for the reaction $H_2O(g + CO\ (g) \Leftrightarrow H\ (g) + CO\ (g)$ is 81 If the rate constant of the forward reaction s 162 lit. m 1^1, sec' what is the velocity constant (in lit. rnole sec) for the backward reaction?

1 3 2) 2 3) 261 4) 243

28) The equilibrium constant for the reaction $H\ (g) + \ (g) \Leftrightarrow 2HI(g)$ is 64 at certain temperature. The equilibrium concentrations of H_2 and HI are 2 and 16 rnol lit^{-1} respectively The equilibrium concentration (in mol Lit^{-1}) of 1 is?

1) 16 2) 4 3) 8 4) 2

29) The equilibrium constant for the reaction $H_2O(l) + CO(g) \Leftrightarrow H_2(g) + CO_2(g)$ is 64. If the rate constant for the forward reaction is 160 rate constant for the backward reaction is

1) 0.4 2) 2.5 3) 6 2 4) $10\ 24\times10^3$

30) If the equilibrium constant for the reaction $2AB \Leftrightarrow A + B$ is 49 What is the value of equilibrium constant for $AB \Leftrightarrow \frac{1}{2}A_2 + \frac{1}{2}B_2$?

1) 49 2) 2401 3) 7 4) 0 02

31) The active mass of 64g of HI in a two litre flask would be

1) 2 2) 1 3) 5 4) 0 25

32) For the reaction $C(s) + CO(g) \Leftrightarrow 2CO(g)$, the partial pressures of CO_2 and CO are 4.0 and 4.0 respectively at equilibrium, what is the value of K for this reaction

1) 0.5 2) 1.0 3) 4.0 4) 32

33) The equilibrium constants for the reactions $N(g) + O_2(g) \Leftrightarrow 2NO(g)$ and $NO(g) + 1/2\ O_2(g) \Leftrightarrow NO_2(g)$ are K_1 and K_2 respectively. Then the equilibrium constant for the reaction $N(g) + 20(g) \Leftrightarrow 2NO(g)$ is

1) $\frac{K_1}{K_2}$ 2) $K_1^2 - K_2^2$ 3) $K_1 \times K_2^2$ $K_1^2 - K_2$

34) The equilibrium constant K of a reversible reaction is 10 The rate constant for the reverse reaction is 2.8 What is the rate constant for the forward reaction

1) 0.28 2) 28 3) 0 028 4) 280

35) The molar concentrations of A, 8 and C at equilibrium for the reaction $A + 28 \Leftrightarrow 3C$ are 2 3 and 4 *molesl* lit respectively. Its K is

1) 2 2) 3 56 3) 0 2 4) 0 026

36) he value of Kp for the reaction $2AB(g \Leftrightarrow 2A(g) + B_2(g)$ is 1.2×10^{-2} at 1065 C The value of K for this reaction is

) $2x10^{-2}$ 2) $<1\ 2x10^{-2}$ 3) 83 4) $> 1.2 \times 10^{-3}$

37) The value of Kp for the reaction $2H\ S(g) \Leftrightarrow 2H_2(g) + S_2(g)$ is $1\ 2 \times 10^{-2}$ at 1065^0 C. The value of K for the reaction is

) 1.2 x 2) $< 1.2 \times 10^{-2}$ 3) 83 4) > 1 2 x 10

38) n the equil brium reaction $N2 + 3H_2 \Leftrightarrow 2NH_3$, the sign of accompanying the reaction is

1) positive 2) negative
3) may be positive or negative 4) cannot be predicted

39) Which of the following is not affected by change in pressure?

1) $N_2(g) + O_2(g) \Leftrightarrow 2NO(g)$ 2) $N(g)\ \ 3H(g) \Leftrightarrow 2\,NH(g)$
3) $PCl;_5(g) \Leftrightarrow PCl\ (g) + Cl$ 4) $2SO_2(g) + O2(g) \Leftrightarrow 2SO(g)$

40. An increase in pressure would favour the following reaction

1) $N_2(g) + O_2(g) \Leftrightarrow 2NO(g)$ 2) $2SO_2(g) + O_2(g) \Leftrightarrow 2SO(g)$
3) $PCl_5(g) \Leftrightarrow PCl_3(g) + Cl_2(g)$ 4) $H(g) + 1(g) \Leftrightarrow 2HI$

41) Hydrolysis of sucrose gives Sucrose + H O $\Leftrightarrow$ Glucos + Fru to e Equilibrium constant K for the reaction is 2 10 at 300K calculate G at 300 K

1) 7 64 10^4 J Mol
2) 7.64 10^{-4} J Mole^{-1}
3) -7.64 x 10^{-4} JMole^{-1}
4) -7.64 x 104 JMole

42) The value of 1G for the phosphorylation of glucose in glycolysis is13.8 kj/mol Find the value of K at 298K

1) 5.8 10^{-5}
2) 5.8 x 10^5
3) 3 8 10
4) 3.8 x 105^{-3}

43) The value of K for the reaction 2A $\Leftrightarrow$ B + C is 2 10^{-3} At a given time, the composition of reaction mixture is [A]=[B]=[C]= 3 10^{-} M In which direction the reaction will proceed?

1) Forward
2) Backward
3) At equilibrium
4) Forward or Backward

44) In the dissociation of HI 20% of HI is dissociated at equilibrium Calculate K_p for HI (g) $\Leftrightarrow$ 1/2 H_2(g) +1/2 (g)

2 2

1) 1.25
2) 0 125
3) 12.5
4) 0 0125

45) For the reaction A + B $\Leftrightarrow$ 3C at 25 A 3 litre vessel contains 1,2 & 4 mole of A,B and C respectively Predict the direction of reaction if: K for the reaction is 10

1) Backward
2) Forward
3) Equilibrium
4) Any direction

46) At equilibrium K =1, then;

1) $\Delta G = 0$
2) ΔG = +ve
3) DG = -ve
4) ΔG = +ve or - ve

47 The equilibrium constant for a reaction + 3H $\Leftrightarrow$ 2NH is K The equilibrium constant for the reaction H $\Leftrightarrow$ 1/2 + 3/2 H_2 is :

1) 1/K
2) K
3) $K^{-1\,2}$
4) K^1

48) At 500 K, the equilibrium constant for the reaction cis- C H Cl $\Leftrightarrow$ trans -C_2H Cl is 0 6 At the same temperature the equilibrium constant for the reaction. trans - C H Cl $\Leftrightarrow$ cis - C H C1

1) 0.6
2) 1 66
3) 1
4) 16.6

49) **Assertion:** A catalyst has no effect on the state of equilibrium

Reason: A catalyst influences the rates of both forward and backward reactions to the same extent.

1) Both A and R are true and R is correct explanation of A.
2) Both A and R are true and R is not correct explanation of A.
3) A is true R is false
4) R is true A is false

50) **Assertion :** Increase of pressure for a reaction which is in equilibrium shifts the equilibrium towards direction where less number of moles are present.

Reason: The change in pressure in an equilibrium react on changes the position or direction of equilibrium when Sn is not equal to zero.

1 Both A and R are true and R is correct explanation of A.

2) Both A and R are true and R is not correct explanation of A.

3) A is true R is false

4) R is true A is false

51) When CO_2 dissolves in water the following equilibrium is established, $CO + 2H_2 \Leftrightarrow H O^{+} + HCC$ - for which the equilibrium constant 3.8×10^{-7} and pH = 6 0 The ratio of [HCO] to [CO] would be

1) 3.8×10^{-1} 2) 3.8 x 10-13 3) 13.4 4) 6.0

52) In a 500 ml flask the degree of dissociation of PCl_5 at equilibrium is 40% and the initial amount is 5 moles The value of equilibrium constant in mol 1t for the decomposition of P 1 is (008)

53) A 550 K, the K, for the following reaction is $^{-4}$ ol^{-1} lit.

+ g) $\Leftrightarrow$ Z(g) At equilibrium it was observed that $[X]=\frac{1}{2}[Y]=\frac{1}{2}[Z]$. What is the value of [Z] (in mol lit^{-1}) at equilibrium?

0^{-} 2) 10^{-4} 3) 2 10^{4} 4) 10

54) f is heated in two seperate vessels of volume 5 lit and 10 lit respectively at 27°C The extent of dissociation of PCl will be

1) More in 5 lit vessel 2) More in 10 lit vessel

3) Equal in both vessels 4) Cannot be said

55) mole of PCl is heated with 0 2 mole of chlorine in a one litre vessel till equilibrium is established At equilibrium if 0 8 mole of chlorine is present in the vessel, the equilibrium constant for the dissociation of PCl is

1) 0 9 2) 1 8 3) 1 2 4) 0 5

56). AB dissociates as AB (g) $\Leftrightarrow$ AB(g) + B(g). When the initial pressure of AB2 is 600mm Hg the total equilibrium pressure is 800mm of Hg Calculate K for the reaction assuming that the volume of the system remains unchanged

1) 50 2) 100 3) 166 8 4) 400

57) 28g of N and 6g of H were mixed. At equilibrium 17g of NH was formed The weight of N and H at equilibrium are respectively

1) 11g & zero 2) 19 & 3g 3)14g&3g 4)11g &3g

58) For the equilibrium A + B $\Leftrightarrow$ C + 0 A and B are mixed in a vessel at T The initial conc of A is twice that of B After equilibrium is reached concentration of C was thrice that of B. Then its K   is

1) 4.5 2) 9 3) 1.8 4) 0.9

59) One mole of A and 2 moles of B are allowed to react in a 0 5 lit flask. What is the value of K_c if at equilibirum 0.4 moles of C is formed in the reaction A + 2B $\Leftrightarrow$C + 20

1) 4/9 2) 9/4 3) 8/27 4) 27/8

60) In the reversible reaction 2NH (g) $\Leftrightarrow$ N_2(g) + 3H (g) when 1 mole of NH is taken initially in a litre flask and 0 2 moles of N are formed at equilibrium its equilibrium constant K is

1) 72 2) 12 10^{-2} 3) 3 4) 27

61 The equilibrium constant for the reaction CH COOH(aq)+ C H OH(aq) $\Leftrightarrow$ CH COOC H (aq) + Hp(l) is 16 The initial concentration of each of CH COOH and C H OH is two (molellit). What is the concentration of CH COOC H at equilibrium in moles / lit

1) 0 9 2) 0 6 3) 1.6 4) 0 7

62 4 5 moles each of hydrogen and iodine are heated in a sealed 10 litre vessel. At equilibrium 3.0 moles of HI were found. The equilibrium constant for

H (g) + 1 (g) $\Leftrightarrow$ 2HI(g) is

1) 1 2) 10 3) 3 4) 0.33

63) HI was heated in a closed tube at 440 C till equilibrium is reaChed. At this temperature 22% of HI was dissociated The equilibrium constant for this dissociation is 1)

0 282 2) 0.0796 3) 0.0199 4) 1 99

64) The equilibrium constant K   for the reaction 2SO + O $\Leftrightarrow$ 2SO is 2 5 atm^{-1} What would be the partial pressure of 0 at equilibrium. If the equilibrium pressures of SO and SO are equal

1) 304 mm 2) 30.4 mm 3) 0 04 mm 4)760 mm

65 A(g) + 28g $\Leftrightarrow$ 2C(g + D(g) was studied using an initial concentration of 8 which was 5 time that of A. But the equilibrium concentration of and 8 were found to be equal. Then K   for the equilibrium is

1) 4 2) 8 3) 6 4) 12

66) 1 mole of H   2 moles of 1   and 3 moles of HI were taken in 1 litre flask The equilibrium constant of the reaction H (g + l_2(g) $\Leftrightarrow$ 2HI(g) is 50 at 330 C. The concentration of HI at equilibrium is (moles 1it)

1) 0 3 2) 1 3 3) 4.4 4) 2 7

67 If 340 g of a mixture of N and H in the correct ratio gave 20   yield of NH   the mass of NH produced would be

1) 16 9 2) 17g 3) 20 9 4) 68 9

68) The Kp of the reaction is NH HS(s) ⇔ NH (g) + H S(g) If the total pressure at equilibrium is 30 atm

1) 15 atm 2) 225atm 3) 30 atm 4) 15 atm

69 In the reaction A + B ⇔ C + D the initial concentration of reactants is 1 mole / lit each The equilibrium concentration of each product is 9 o es. lit . Its K is 0 9 2 0 81 3) 81 4) 8 1

70) e o e of H g is heated with one mole of 1 g was found that 1.48 moles of HI(g) is o ed at equilibrium Its K is

1) 16 2 32 3) 8 4) 24

KEY

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

7b. ACIDS AND BASES

Synopsis:

1. Acids are substances which 1) are sour to taste and corrosive, 2) turn blue litmus red, 3) contain Hydrogen part which can be displaced by metals. 4) reacts with bases to form salts and 5) react with carbonates and bicarbonates liberating CO_2.
2. Bases are substances which 1) are bitter to taste, 2) turn red litmus blue, 3) contain OH- part which can be displaced by non-metals. 4) Reacts with acids to form salt and 5) they are soapy to touch.
3. According to Arrhenius an acid ionizes in solution producing Hydrogen ions and a base ionizes in solution producing hydroxyl ions.
4. The degree of ionization of an acid or a base is a measure of its strength.
5. Since strong acids and strong base ionize to a greater extent in water, their aqueous solutions are good conductors.
6. According to Arrhenius neutralisation is the reaction between hydrogen ions and Hydroxyl ions to form water.
   $H^+ + OH^- \rightarrow H_2O$
7. Arrhenius theory of acids and bases explains the behaviour of acids and bases in aqueous solutions only. It fails to explain the behaviour of acids and bases in non-aqueous media.
8. According to Lowry and Bronsted an acid is a proton donor and a base is a proton acceptor.
9. The ease or readiness with which an acid loses protons is a measure of its strength.
   The ease or readiness with which a base gains protons is a measure of its strength.
10. For an acid or a base to exhibit its character there should be a proper medium.
11. Proton accepting (or basic) solvents like water, ammonia, ethyl alcohol etc, are called protophilic solvents.
12. Solvents like water which can accept as well as donate protons are called amphiprotic solvents.
13. Solvents like Benzene, Carbon tetrachloride etc, which can neither accept nor donate protons are called aprotic solvents.
14. Even a strong acid does not exhibit its acidic character in the absence of any medium or in an aprotic medium like benzene.
15. Pure and dry HCl is not an acid. It does not turn blue litmus red. Similarly a solution of HCl in Benzene is not an acid.
16. Even a weak acid like acetic acid behaves as a strong acid in strongly protophilic medium like liquid ammonia.
17. Similarly even a weak base like aniline behaves as a strong base in strongly protogenic or acidic medium like liquid HF.
18. Strong acids as well as weak acids behave as equally strong acids in a basic medium like liquid ammonia. This is known as leveling effect.
19. To compare the strength of strong acids, the comparison is made in a weakly protophilic medium like acetic acid.
20. Even a strong acid behaves like a weak acid in a weak protophilic medium like acetic acid and the solution is a poor conductor.

21. When the strengths of strong acids are compared in acetic acid, their strengths are found to be in the order.

$$HClO_4 > HI > HBr > H_2SO_4 > HCl > HNO_3.$$

22. The strongest acid is perchloric acid $HClO_4$.
23. In the case of oxyacids of a family the acidic strength decreases with an increase in the atomic weight of the central atom.

    Ex: In VI group acidic strengths are in the order $H_2SO_4 > H_2SeO_4 > H_6TeO_6$.
24. Acid-Proton = Conjugate base
25. Acid +Proton = Conjugate acid
26. If the acid is strong its conjugate base is a weak and vice versa.
27. Ions like HCO_3^-, HSO_4^-, $H_2PO_4^-$, etc--, can lose and also gain protons. Hence they can behave as Bronsted acids as well as Bronsted bases.
28. The higher the value of K_a the greater is the strength of the acid.
29. To indicate the strength of an acid we can also use the p^{ka} value. $p^{ka} = -\log Ka$. The higher the value of Ka the smaller will be the value of P^{ka}.
30. An acid with a smaller p^{ka} value is stronger than an acid with a higher p^{ka} value.
31. If the ionization constants of two weak acids are Ka_1 and Ka_2, then

$$\frac{\text{Strength of acid}_1}{\text{Strength of acid}_2} = \sqrt{\frac{Ka_1}{Ka_2}}$$

32. When a base is dissolved in water, the following equilibrium exists.

$$B + H_2O \Leftrightarrow BH^+ + OH^-$$

$$K_b = \frac{C_{BH} + C_{OH}^-}{C_B}$$

    K_b is called the dissociation constant or ionization constant of the base.
33. In Arrhenius theory the ionization of a base is represented as $BOH \Leftrightarrow B^+ + OH^-$

$$K_b = \frac{C_{BH} + C_{OH}^-}{C_{BOH}}$$

34. If the dissociation constants of two weak bases are Kb_1 and Kb_2 then

$$\frac{\text{Strength of base}_1}{\text{Strength of base}_2} = \sqrt{\frac{Kb_1}{Kb_2}}$$

35. Limitations of Bronsted and Lowry theory:
    a. Proton donations and acceptance happens only in the presence of another substance.
    b. It failed to explain the acidic nature of electron deficient molecules like BCl_3, $AlCl_3$, BF_3 etc.

    Advantages of Bronsted and Lowery theory:
    a) This theory explains the behaviour of acids and bases in both aqueous and non aqueous solvents.
    b) It explains the behaviour of NH_3 and CaO as bases.
    c) This is a general theory accounting for large number of acids and bases than Arrhenius theory.
36. Lewis theory of acid and bases:

    According to this theory:
    a. A base is an electron pair donor or a nucleophile and
    b. An acid is an electron pair acceptor or an electrophile.

37. a. A molecule or ion which contain a lone pair of electrons can act as a Lewis bases.

b. All anions like Cl^-, Br^-, I^-, NH_2^-, F^-, SCN^-, OH^- are Lewis bases.

c. Molecules with multiple bonds like, CO, NO, $CH\equiv CH$, $CH_2=CH_2$.

38. a. A molecule or ion which contains a vacant orbital in its valence shell can act as a Lewis acid.

b. Molecules whose central atom has an incomplete octet with an empty orbital like BF_3, BCl_3, BBr_3, $AlCl_3$ are lewis acids.

c. Elements with electron sexet.
Ex: O and S

39. According to Lewis theory neutralization results in the formation of a coordinate covalent bond between an acid and a base.

40. Limitation of Lewis theory:

a. This theory failed to explain the strength, of acids and bases.

b. Acid-base reactions are instantaneous. But Lewis acid-base reactions are very slow.

c. All acid-base reactions do not involve dative bond formation.

d. The catalytic activity of H^+ ion was not explained by this theory.

41. Ionic product of water $[K_w]$:

a. Water undergoes self-ionisation and in water or in any aqueous solution the following equilibrium exists

$$2H_2O \Leftrightarrow H_3O^+ + OH^-$$

b. In water or in any aqueous solution the product $C_H^+ . C_{OH}^-$ is a constant at a constant temperature and is known as the ionic product of water K_w.

c. $Kw = C_H^+ . C_{OH}^- = 10^{-14}$ mole^2/lit^2 at 25^0C.

42. Water is a stronger acid than D_2O because K_w for H_2O (10^{-14}mol^2 / lt^2) is greater than K_w for D_2O which is $0.3x10^{-15}$ mol^2 / lt^2.

43. a. With an increase of temperature ionization of water increases and so K_w increases.

b. With increase in temperature of $[H^+]$ increases.

c. With increase in temperature $[OH^-]$ increases.

d. With increase in temperature p^H or p^{OH} decreases.

44. In pure water or in a neutral solution.

$$C_H^+ = C_{OH}^- = 10^{-7} \text{ mole/lit at } 25^0C.$$

45. In an acidic solution $C_H^+ > C_{OH}^-$

or $C_H^+ > 10^{-7}$ mole/lit and or $C_{OH}^- > 10^{-7}$ mole/lit at 25^0C.

46. In an alkaline solution $C_H^+ < C_{OH}^-$ or $C_H^+ < 10^{-7}$ and $C_{OH}^- > 10^{-7}$ mole/lit 25^0C.

p^H of a Solution:

47. S.P.L Sorensen devised the p^H scale. (p^H means puissance de hydrogen)

$$p^H = -\log[H^+] = \log \frac{1}{[H^+]}$$

48. a. In similar way $p^{OH} = -\log C_{OH}^-$

b. $P^{Ka} = -\log_{10}[K_a]$

c. $P^{Kb} = -\log_{10}[K_b]$

d. $P^{Kw} = -\log_{10}[K_w]$

49. a. For any solution $p^H + p^{OH} = 14$ at 25^0C.

b. $K_a \times K_b = 10^{-14}$ [For a conjugate acid base pair at 25^0C]

c. $p^{Ka} + p^{Kb} = 14$ [For a conjugate acid base pair at 25^0C]

d. $p^{Ka} + p^{Kb} = p^{Kw}$ at any temperature [For a conjugate acid base pair]

50. The p^H of 10^{-8} M HCl is 6.96.
51. The p^H of 10^{-8} M HCl is 7.04.
52. However dilute an acidic solution may be its p^H cannot be greater than 7. It must be less than 7.

Buffer solution:

53. Solution whose p^H is a constant and whose p^H does not alter on the addition of a small quantities of a strong acid or a strong base or any amount of water is called a buffer solution.
54. Buffer solutions are also known as solutions of reserve acidity and alkalinity.
55. a. A buffer solution is generally a mixture of a weak acid and its salt with a strong base.

    b. Excess of weak acid with a strong base.

    Ex: 1) $CH_3COOH + CH_3COONa$

    2) 2M CH_3COOH + 1M NaOH
56. Other Examples are Boric acid + Borax, Phosphoric acid + Ammonium Phosphate etc, are acidic buffers.
57. Ammonium acetate solution can act as a neutral buffer solution.
58. A buffer solution can also be defined as a mixture of a weak acid and its conjugate base or a mixture of weak base and its conjugate acid.
59. a. In a mixture of CH_3COOH+CH_3COONa, reserve acidic nature is due to CH_3COOH and reserve alkalinity is due to CH_3COO^-.

    b. To this buffer when some H+ ions are added the reaction taking place is $CH_3COO^- + H^- \rightarrow CH_3COOH$, Hence the p^H remains constant.

    c. To this buffer when some OH^- ions are added, the reaction taking place is $OH^- + CH_3COOH \rightarrow CH_3COO^- + H_2O$. Hence the p^H remains constant.
60. a. In a mixture of NH_3+NH_4Cl, reserve acidic nature is due to the presence of NH_4^+ ions and reserve alkalinity is due to the presence of NH_3 [orNH_4OH]

    b. To this buffer if some OH^- ions are added then the reaction is $NH_4^+ + OH^- \rightarrow NH_4OH$. NH_4OH is a weak base, hence the pH remains constant.

    c. To this buffer if some H^+ ions are added, then the reaction is $NH_4OH + H^+ \rightarrow NH_4^+ + H_2O$. Hence the p^H remains constant.
61. Blood is an acidic buffer solution. The p^H value of blood is 7.4.
62. The p^H of an acidic buffer is given by Henderson equation.

    $p^H = p^{Ka} + \log C_{salt} / C_{acid}$
63. The p^{OH} of a basic buffer is given by

    $P^{OH} = p^{Kb} + \log C_{salt} / C_{base}$

    $$\therefore p^H \text{ of basic buffer} = p^{Kw} - \left(p^{Kb} + \log \frac{C_{salt}}{C_{base}} \right) = 14 - p^{Kb} - \log \frac{C_{salt}}{C_{base}}$$
64. Buffer capacity (ϕ) is defined as the number of moles of acid or base added per litre of buffer solution to produce unit change in the p^H.
65. A buffer solution has maximum buffer capacity when its $p^H = p^{Ka}$ or when $C_{salt} / C_{acid} = 1$ or when C salt = C acid.

66. A buffer solution is satisfactory when its p^H is with in the range p^{Ka} –1 to p^{Kb}+1. In the case of basic buffer when its p^{OH} is within the range p^{Ka} –1 to p^{Kb}+1 it is satisfactory.

67. Buffer solutions are useful in 1) Chemical analysis, 2) Industrial reactions, 3) Enzyme catalysed reactions and 4) Biochemical processes 5) Buffer solutions are used in water softening.

68. Salts of strong acids and strong bases like NaCl, KCl etc, do not undergo hydrolysis. Their aqueous solutions are neutral, with p^H =7.

69. a. Salts of weak acids and strong bases like CH_3COONa, $Na_2\,CO_3$, KCN, Na_2S etc, undergo anionic hydrolysis. Their aqueous solutions are alkaline. The p^H of their solutions is greater than 7.

    b. $A^-+H_2O \Rightarrow HA+OH^-$

    c. For such salts the hydrolysis constant

$$K_h = \frac{K_w}{K_a}$$

70. Salts of weak bases and strong acids like NH_4Cl, $CuSO_4$, $FeCl_3$, $MgCl_2$, $CaSO_4$ etc, undergo cationic hydrolysis. Their aqueous solutions are acidic. The pH of the solution is less than 7.

    a. $B^+ +H_2O \Rightarrow BOH + H^+$ or $B^+ +H_2O \Rightarrow B+H_3O^+$

    b. For such salts the hydrolysis constant.

$$K_h = \frac{K_w}{K_b}$$

71. a) Salts of weak acids and weak bases like CH_3COONH_4 etc, undergo both cationic and anionic. In the case of ammonium acetate, the acid and base are equally strong. Hence ammonium acetate solution is neutral. If the acid is weaker as in the case of $(NH_4)_2CO_3$ and NH_4CN, then the solution is slightly alkaline. If the base is weaker, the solution is slightly acidic. In Ammonium phosphate, the solution is slightly acidic.

    b) The nature of such a solution depends on relative values of K_a and K_b.

    If $K_a = K_b$ -----------Neutral
    If $K_a>K_b$------------Acidic
    If $K_a<K_b$-----------Basic

Acid base indicators

72. The external chemical added to the contents in a conical flask during titration in the visual detection method of end point by a colour change is known as an Indicator.

73. An indicator is a substance which changes its colour with a change in the p^H of the solution.

74. Generally an indicator has one colour in acid medium and another colour in alkaline medium.

75. Indicators used in redox titration are called acid-base Indicator.
    Ex: Phenolpthalein, Methyl orange.

76. The p^H range of an indicator is the p^H range where a detectable visual colour change in indicator is noticed.

77. In general the p^H range of an indicator is P^{Ki} –1 to P^{Ki} +1

78. The p^H range of titration is the change in the p^H near the end point.

79.

Indicator	p^H range	Colour change
Thymol blue (acid)	1.2 to 2.8	Red to Yellow
Methyl orange	3.1 to 4.4	Red to Yellow
Methyl red	4.2 to 6.2	Red to Yellow
Litmus	5.5 to 7.5	Red to Blue
Phenol red	6.8 to 8.4	Yellow to Red
Thymol blue (base)	8.0 to 9.6	Yellow to Blue
Phenolphthalein	8.3 to 10	colourless to Pink

80. **Universal indicator**: The indicator that covers a wide range of p^H (3 to 11) and gives different colours at difference p^H values is called universal indicator.
It is not used in acid-base titrations.
It is prepared by dissolving 0.1g of phenolpthalein 0.2g of Methyl red, 0.3g of Methyl yellow, 0.4g of Bromo thymol blue, 0.5g of thymol blue in 500ml of absolute alcohol and a little NaOH is added till it acquires a pale yellow colour.
Universal indicator gives red colour at $p^H = 2$, orange colour at $p^H = 4$, yellow colour at $p^H = 6$ green colour at $p^H = 8$ and Blue colour at $p^H = 10$.

Solubility: The no. of grams of solute dissolved in 100gm of solvent is knows as solutbility.

$$\text{Solubility} = \frac{\boldsymbol{weight\ of\ solute}}{\boldsymbol{weight\ of\ solute}} \times \mathbf{100}$$

Units: moles / litre

Solubility depends on following factors:

a) Nature of the solute
b) Nature of solvent
c) Temperature
d) Pressure

Lighter salts dissolve more than heavier salts.

Based on solubility, types of salts.

Soluble salts: which contain more than 1 gm of solute in 100 ml solution.

Eg: NaCl, KCl, $BaCl_2$ etc.

Sparingly soluble salts: Which contains less than 1gm of solute in 100ml solution

Eg: AgCl, AgI

Insoluble salts: Which contain 10^{-3} gm of solute in 100 ml of solution.

Based on amount of dissolved solute in solution, solutions are three types.

Saturated solution: The solution which contain exactly the required amount of solute for saturation.

- It is a stable solution.
- In this solution equilibrium state is established between undissolved solute and dissolved solute.
- It gives unsaturated solution on heating and gives saturated solution on cooling.
- A saturated solution contains.
  a) salt in the solid form b) salt in the dissolved form
- Saturated solution contains two dynamic equillibric. They are
  Un dissolved salt = dissolved salt

Dissolved salt = ions of the salt.
So it is respresented as

AB	=	AB	=	A^++B
(undissolved)		(dissolved)		(ions)

Unsaturated solution: The solution which contains lesser amount of solute that required for saturation.

Supersaturated solution: The solution which contain which contains more amount of solute that require for saturation.

Solubility product: KSp

- It is the produt of the molar concentration of its ions in a saturated solution of electrolyte.
- KSp for a salt changes with temperature.
- Solubility product of a sparingly soluble salt

$$A_xB_y = xA^{+y} + yB^{-x}$$

$$KS_P = [A^{+y}]^x \quad [B^{-x}]^y$$

Applications: The concept of solubility product helps the formation of a precipitate.

a) ionic product < KSp – no pdt is formed (Unsaturated solution)
b) ionic product >KSp – ppt is formed (Super saturated solution)
c) ionic product = KSp saturated solution.

- In predicting the solubility of a sparingly soluble salt.
- In quantitive analysis the separation and identification of various basic radicals in different groups.
- In purification of common salt.
- In salting out of soap.
- For binary salts like AgCl, ABr, solubility 'S' can be calculated as $S = \sqrt{K}$ for tertiary salt like $PbCl_2$

$$S = \sqrt{\frac{KS_p}{4}}$$ For AB salt KSp = S

AB_3 salt KSp = $27S^3$
AB_2 salt KSp = $4S^3$

The suppression of the dissociation of a weak electrolyte by addition of strong electrolyte having common ion.

Eg: A solution of sodium acetate is added to solution of acetic acid as a result in the decreases in ionization of acetic acid due to common ion is CH_3COO^-

CH_3COONa = CH3COO$^-$ + 4^+
CH_3COONa = CH3COO$^-$ + Na^+

Accordingly to lechatelier principle [N^+] decreases P^H increases and equilibrium shifts to left hand side.

Applications:

It is is used in controlling H^+ ion concentration in buffer solution.

The concentration of S^{-2} in the II group. VOH- in IIIgP of qualitative analysis is controlled by HCl and NH_4OH on the basis of common in effect.

QUESTIONS BANK

1) Arrhenius neutralisation involves

1) formation of dative bond
2) formation of water by the combination of H^+ with OH^-
3) transfer of proton   4) formation of H_3O^+

2) NH_3 is not a base according to

1) Bronsted theory   2) Lewis theory
3) Arrhenius theory   4) Lowry theory

3) HCI gas cannot act as an acid in

1) Polar solvents   2) highly polar solvents
3) NH_3 liquid   4) non-polar solvents.

4) Which of the following is Bronsted base

1) H_3O^+ 2) NO_3^- 3) NH_4^+ 4) CH_3COOH

5) The conjugate acid of water is

1) OH^- 2) H^+ 3) H_3O^+ 4) H_3O^-

6) HCl does not behave as acid in

1) Water   2) Ammonia   3) Benzene   4) Aq NaOH

7) Which of the following is the conjugate base of $[C_2H_5NH_3]^+$?

1) $[C_2H_5N_H]^-$ 2) $[C_2H_5NH_3]OH$ 3) $C_2H_5NH_2$ 4) $C_2H_5NH_2^+$

8) The strongest base of the following species is

1) NH_2^- 2) OH^- 3) $O2^-$ 4) S^{2-}

9) The conjugate base of H_2SO_4 in the following reaction is

$$H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-$$

1) H_2O 2) $HSO4^-$ 3) $H3O^+$ 4) SO_4^{-2}

10) The following equilibrium is established whe hydrogen chloride is dissolved in acetic acid . $HCl + CH_3COOH \rightleftharpoons CI^- + CH_3COOH_2^+$

The set that characterizes the conjugate acid - base pairs is

1) (HCI, CH_3COOH) and $(CH_3COOH_2^+, CI^-)$
2) (HCI, CH_3COOH) and (CH_3COOH, CI^-)
3) (CH_3COOH2^+, HCI) and (CI^-, CH_3COOH)
4) (HCI, CI^-) and $(CH_3COOH_2^+, CH_3COOH)$

11) In the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

1) NH_3 is the conjugate base of the acid OH^-
2) H_2O is the conjugate acid of the base NH_3
3) NH_3 is the conjugate acid of the base NH^+
4) NH_3 is the conjugate base of the NH_4^+

12) Which anion is weakest base?

1) C_2H5O^- 2) NO_3^- 3) F^- 4) CH_3COO^-

13). The strongest Bronsted base among the following is

1) ClO^- 2) ClO_2^- 3) ClO_3^- 4) ClO_4^-

14) Identify Bronsted - Lowry acids in the reaction given.

$$[Al(H_2O_6)^{+3} + HCO_3^- \rightleftharpoons [Al(H_2O)_5(OH^-)]^{2+} + H_2CO_3$$

A B C D

The correct answers

1) A, C 2) B, D 3) A, D 4) B, C

15) The conjugate acid of SO^{-2}_4 is

1) HSO_4^- 2) H_2SO_4 3) HSO_3^{-2} 4) HSO_3^{-2}

16) Buffer Solution is prepared by mixing

.1) Weak acid and its salt of strong base.
2) Strong acid + its salt of strong base
3) Weak acid + its salt of weak base
4) Strong base + its salts of strong acid

17) **Assertion A:** P^H of a buffer solution does not change on dilution.

Reason R: On dilution the ratio of concentration of salt and acid (or base) remains unchanged. hydrolysis

1) 'A' and 'R' are true, 'R' is correct explanation of 'A'.
2) 'A' and 'R' are true. 'R' is not correct explanation of 'A'.
3) 'A' is true but 'R' is false.
4) 'A' is false but 'R' is true.
5) Both 'N and 'R' are false.

18) An acidic buffer has the maximum buffer capacity when

1) [Acid] > [Salt] 2) [Acid] = [Salt]
3) [Acid] < [Salt] 4) [Acid] = 2x[Salt]

19) When CH_3COONa is added to an aqueous solution of CH_3COOH

1) The pH of the solution decreases
2) The pH of the solution increases
3) The pH of the solution remains unaltered
4) An acid salt is produced

20) In aqueous solution, the following mixture acts as buffer

1) $HNO_3 + KNO_3$ 2) $H_2SO_4 + K_2SO_4$
3) $NH_4OH + NH_4Cl$ 4) $CH_3COOH + NaCl$

21) The strength of acid is highest in

1) $pK_a = 6$ 2) $pK_a = 5$ 3) $pK_a = 10$ 4) $K_a = 10 - 11$

22) The dissociation constants of two acids HA_1 and HA_2 are 4.5×10^{-4} and 1.8×10^{-5} respectively. If both are having equal concentrations the relative strength of acids is

1) 5:2 2) 2:5 3) 5:1 4) 1:5

23) The ionic product of water $[H^+]$ $[OH^-]$ at 25°C is.mole^2/lit^2

1) 10^{-7} 2) 10^{-12} 3) 10^{-14} 4) 10^{-13}

24) At 70°C the concentration of H^+ ion in aqueous solution of NaCl is 10^{-6} mole/lit. The OH^- ion concentration is 254

1) 10^{-8} moles/lit. 2) 10^{-6}moles/lit. 3) 10^{-7} moles/lit. 4) 10^{-9}moles/lit.

25) At 90^0 C for pure water $[H_3O^+] = 10^{-6.5}$moles/lit. The value of K_w at this Temperature is

1) 10^{-6} 2) 10^{-13} 3) 10^{-14} 4) 10^{-8}

26) The hydronium ion concentration of a solution is 1×10^{-4} M at 25^0C. The hydroxide Ion concentration in the same solution is

1) 1×10^{-14} M 2) 1×10^{-7} M 3) 1×10^{-10} M 4) 1×10^{-4} M

27) A 80°C, distilled water has hydronium ion (H_3O^+) concentration equal to 1×10^{-6} mol l^{-1}. The value of K_w at this temperature would be

1) 1×10^{-6} 2) 1×10^{-12} 3) 1×10^{-9} 4) 1×10^{-14}

28) The ionic product of water is 10^{-14},What is the H^+ ion concentration of a 0.01 M NaOH solution

1)’10^{-14} M 2) 14 M 3) 13 M 4) 10^{-12} M

29) The Ka of a weak acid is 10^{-5}, pKb value of its conjugate base is

1) 7 2) 5 3) 9 4) 6

30) In. 0.001 M HNO_3, OH^- ion concentration is

1) 1×10^{-3} 2) 1×10^{-11} 3) 1×10^{-14} 4) 1×10^{-3}

31) pH of one litre solution containing 40gm of NaOH

1) 2 2) 10 3) 8 4) 14

32) The pH of 0.002 M - H_2SO_4 solution is

1) 2 2) 2.3979 3) 2.6021 4) 2.699

33) If a solution of pH = 5 is diluted 100 times, its pH Becomes

1) 7.3010 2) 7 3) 6.699 4) 6.3010

34) The pH of NaOH solution is 12. What is the amount in grams of NaOH present in one litre of a solution?

1) 40 2) 4 3) 0.4 4) 20

35) The pH of a solution at 25°C is 5. The $[OH^-]$ of the solution is

1) 10^{-5} M 2) 1×10^{-9} M 3) 10^{-7} M 4) 10^{-8} M

36) The hydrolysis constant K_h of a salt of NaOH and a weak acid (HX) if the K_a of the acid is 2×10^{-6} is

1) 5×10^{-8} 2) 5×10^{-6} 3) 5×10^{-9} 4) 2.5×10^{-7}

37) Hydrolysis constant (K_h) of ammonium acetate ($K_a = 1.8 \times 10^{-5}$, $K_b = 2 \times 10^{-5}$, $K_w = 10^{-14}$) is

1) 28×10^{-5} 2) 2.8×10^{-5} 3) 0.28×10^{-10} 4) 1.8×10^{-5}

38) Calculate the hydrolysis constant of a salt of weak acid ($K_a = 2 \times 10^{-6}$) and of a weak base ($K_b = 5 \times 10^{-7}$)

1) 10^{-4} 2) 10^{-2} 3) 10^{-6} 4) 10^{-8}

39) One litre of a buffer solution contains 0.01M NH_4Cl and 0.1M NH_4OH. The p^{kb} of base is 5. pH value of the solution is

1) 9 2) 10 3) 4 4) 6

40) 5 moles of acid is required to change the pH of 1 litre of buffer by 2 units. Find buffer capacity.

1) 5 2) 3.5 3) 2.5 4) 15

41) The solubility of calomel in water at 25°C is x moles/lit. Its solubility product is

1) $4x^3$ 2) $12x^3$ 3) $108x^5$ 4) x^2

42) Which is correct about ionic product of an electrolyte?

1) It is always equal to its solubility product
2) Can be < (or) equal to K_{SP}
3) Always < K_{SP}
4) Can be < (or) equal to (or) > K_{SP}

43) The solubility product of $NaCl$ is 2.5×10^{-9} moles^2 / lit^2. Its solubility is

1) 2.5×10^{-9} mol/lit 2) 5×10^{-5} mol^2 / lit^2
3) 5×10^{-5} mol/lit 4) 1.3×10^{-5} mol/lit

44) Amongst the following hydroxides, the one which has the highest value of K_{SP} at 25° C

1) KOH 2) CsOH 3) $LiOH$ 4) $RbOH$

45) K_{SP} of $BaCl_2$ is

1) $4S^2$ 2) $5S^3$ 3) $4\,S^3$ 4) None of these

46) The solubility of salt xy_2 is

1) $4\,K_{SP}$ 2) $\frac{K_S}{4}$ 3) $\sqrt[3]{\frac{K_{SP}}{4}}$ 4) $\sqrt{\frac{K_{SP}}{4}}$

47) The solubility of Aluminium sulphate is 1.0×10^{-4} *mol/lit*, the solubility product is

1) 108×10^{-4} 2) 10800×10^{-20}
3) 1080×10^{-18} 4) 1.08×10^{-18}

48) The solubility of $AgCl$ in $NaCl$ solution is less than that in pure water, because

1) Solubility product of $AgCl$ is less than of $NaCl$
2) Common ion effect
3) Both 1 and 2 4) None of these

49) In 0.1 M- solution, a mono - basic acid is 1% ionized. The ionisation constant of the acid is

1) 1×10^{-3} 2) 1×10^{-7} 3) 1×10^{-5} 4) 1×10^{-14}

50) In 0.01 M solution, HCN is 0.01% ionized. Its Ka is

1) 10^{-6} 2) 10^{-10} 3) 10^{-3} 4) 0.02

51) A solution of an acid has $P^H = 4.70$. Find out the number of OH^- ions ($P^{kW} = 14$)

1) 5×10^{-10} 2) 4×10^{-10} 3) 2×10^{-5} 4) 9×10^{-4}

52) The hydrogen concentration of 0.01 N NaOH solution is

1) 1×10^{-2} mol/L 2) 10^{-14} mol/L 3) 10^{-12} mol/L 4) 10^{-13} mol/L

53) Calculate the pH value in 10^{-8} M NaOH

1) 7.04 2) 8 3) 6 4) 6.96

54) when 100ml of M/10 NaOH solution and 50 ml of M/5 HCl solution are mixed, the pH of the resulting solution would be

1) 0 2) 7 3) < 7 4) > 7

55) pH of water is 7. When substance Y is dissolved in water, the pH becomes 13. The substance Y is salt of

1) Strong acid and strong base 2) Weak acid and weak base
3) Strong acid and weak base 4) Weak acid and strong base.

56) The pH of 0.0001 M NaOH is
1) 4 2) 10 3) 12 4) 11

57) How much sodium acetate should be added to 0.1M solution of CH_3COOH to give a solution of pH 5.5 (pka of $CH_3COOH = 4.5$)?
1) 0.1M 2) 0.2M 3) 1.0M 4) 10.0M

58) In a mixture of weak acid and its salt, the ratio of concentration of acid to salt is increased ten fold. The pH of the solution is
1) Decreases by 1 2) Decreases by 1/10
3) Increases by 1 4) Increases ten fold.

59) The pH of the solution is the negative logarithm to the base 10 of its hydrogen ion concentration in
1) Moles per litre 2) Milli moles per litre
3) Micro moles per litre 4) Nano moles 1 per litre.

60) If pH of A, B, C and Dare 9.5, 2.5, 3.5 and 5.5 respectively, the strongest acid is
1) A 2) B 3) C 4) D

61) Degree of dissociation of CH_3COOH and NH_4OH are the same. If 0.01 M solution of CH_3COOH has pH=4.0; then pH of 0.01 M NH_4OH will be
1) 4 2) 7 3) 10 4) 14

62) The pH value of 0.1 mol/litre HCl is approximately 1. The approximate pH value of 0.05 mol/ litre H_2SO_4 is most likely to be
1) 0.05 2) 2 3) 0.5 4) 1

63) The pH of pure water at 25^0C is close to
1) 0 2) 7 3) 2 4) 9

64) 100ml of 0.15M HCl is mixed with 100 ml of 0.005M HCl, what is the PH of the resulting solution approximately
1) 2.5 2) 1.5 3) 2 4) 1

65) 50 litres of 0.1M HCl is thoroughly mixed with 50 litres of 0.2M NaOH The pOH of the resulting solution is
1) 1.0 2) 0.0 3) 7.0 4) 1.30

66) The pH of.a solution is increased from 3 to 6. Its H+ ion concentration will be
1) Reduced to half 2) Doubled
3) Reduced by 1000 times 4) Increased by 1000 times

67) The pH of an aqueous solution of H_2O_2 is 6.0. Some chlorine is bubbled through this solution. The pH of the resulting solution will be
1) 6 2) 7 3) Less than 6 4) More than 7

68) The pH of NaOH solution is 12. What is the amount in grams of NaOH present in one litre of a solution?
1) 40 2) 4 3) 0.4 4) 20

69) 75ml of 0.2M HCl is mixed with 25ml of 1M HCl. To this solution 300ml of distilled water is added. What is the pH of the resultant solution:
1) 1 2) 2 3) 4 4) 0.2

70) The pH of 0.1M solution of CH_3COOH if it ionizes to an extent of 1% is.
1) 1 2) 2 3) 3 4) 4

71) 100 ml of 0.3N-Acetic acid solution is mixed with same volume of 0.2 N sodium hydroxide solution. Ionisation constant of acetic acid is 2x10 ^{5} The pH of the mixture is

1) 10 2) 7 3) 5 4) 9

72) The pka of HCN is 9.3.The pH of a solution prepared by mixing 2.5 moles of HCN and 0.25 moles of KCN in water and making of the total volume to 500 ml is

1) 9.3 2) 7.3 3)10.3 4) 8.3

73) When the following is added to 20ml of 0.1 Ml CH_3COOH+20 ml of CH_3COONa. The pH of the! solution does not change

1)10ml of 0.1M CH_3COOH 2)10ml of 0.1M CH_3COONa

3) 10ml of water 4) 20ml of 0.2M CH_3COOH

74) Which of the following mixture is not a buffer solution

1) 100ml of 0.5NHCI and 100 ml of 0.6N-NH_4OH

2) 100 Ml of 0.6 N - HCN and 100ml of 0.4 N-NaOH

3) 100ml of 0.2N NH_4OH and 100ml of 0.2N-CH_3COOH

4) 100ml of 0.4N-HCI and 100 ml of 0.4N -NaOH

75) The ionisation constant of acetic acid is 2 x 10^{-5} The pH of buffer containing acetic acid and sodium acetate is 4.7. The acid and salt ratio in the buffer is.

1) one : one 2) Ten : one 3) one : ten 4) two : one

76) The solubility of Ag_2CrO_4 is 2 x 10^{-2} mol/lit its solubility product is

1) 3.2 x 10^{-5} 2) 32 x 10^{-8} 3) 16 x 10^{-8} 4) 3.32 x 10^{-10}

77) Solubility of AgCl will be minimum in

1) 0.01M Na_2SO_4 2) 0.01 M $CaCl_2$ 3) Pure water 4) O.OOlM $AgNO_3$

78) Solubility product of $Ba(OH)_2$ and $Al(OH)_3$ are 1.8 x 10^{-10} and 2.4 x 10^{-20} respectively. If both Al^{3+} and Ba^{+2} ions are present is a solution, which one will be ppt on addition of Ammoniumn hydroxide solution

1) $Ba(OH)_2$ 2) $Al(OH)_3$

3) Both are precipitated at same time 4) Both are not precipitated

79) What would be the solubility of silver chloride in 0.10 Molar NaCI solution. (Ksp for AgCl solution= 1.20x10^{-10})

1) 0.1 M 2) 1.2x10^{-9} 3) 1.2x10^{-6} 4) 1.2 x 10^{-10}

80) NH_4OH is weak base but it becomes still weaker in the aqueous solution of

1) 0.1MHCI 2) 0.1MNH_ACl 3) 0AMH_2SO_4 4) 0.1MCH_3COOH

KEY

1.2	2.3	3.4	4.2	5.3	6.3	7.3	8.1
9.2	10.4	11.4	12.2	13.1	14.3	15.1	16.1
17.1	18.2	19.2	20.3	21.2	22.3	23.3	24.2
25.2	26.3	27.2	28.4	29.3	30.2	31.4	32.2
33.3	34.3	35.2	36.3	37.2	38.2	39.2	40.3
41.1	42.4	43.3	44.2	45.3	46.3	47.4	48.2
49.3	50.2	51.1	52.3	53.1	54.2	55.4	56.2
57.3	58.1	59.1	60.2	61.3	62.4	63.2	64.2
65.4	66.3	67.3	68.3	69.1	70.3	71.3	72.4
73.3	74.4	75.1	76.1	77.2	78.2	79.2	80.2

8. HYDROGEN AND ITS COMPOUNDS

Synopsis:

1. The first and the lightest element in the periodic table is hydrogen.
2. Element with no neutron is hydrogen.
3. Hydrogen is inflammable gas. Most abundant element in outer atmosphere of the sun (90%).
4. Hydrogen is most abundant element in the universe (92%) and ninth abundant in earths crust.
5. Isotopes of hydrogen $_1H^1$ = protium = light hydrogen.
   $_1H^2$ or D = Deuterium = heavy hydrogen.
   $_1H^3$ or T= tritium = Radioactive hydrogen.
6. Number of neutrons in protium, deuterium & tritium are 0, 1 and 2 respectively. The three isotopes of hydrogen differ in their physical properties.
7. Ordinary hydrogen contains 99.98% protium & 0.02% deuterium.
8. Hydrogen molecule is diatomic (H_2). It occur in two different forms, known as ortho hydrogen and para hydrogen ortho & para hydrogen are spin isomers. In ortho hydrogen molecular spins of H-atoms are in same direction. In para hydrogen molecule the nuclear spins of two atoms are in opposite directions.
9. At absolute zero, H_2 gas contains 100% para form of the molecules. Ordinary hydrogen contains 75% ortho form and 25% para form of the molecules.
10. Atomic hydrogen (nascent hydrogen) is highly reactive and best reducing agent.
11. Uses of hydrogen :-
    (a) In preparation of NH_3, HCl, methyl alcohol.
    (b) In hardening of oils, industrial fuel, in fuel cells.
    (c) As reducing agent in metallurgy.
    (d) Oxy-atomic hydrogen flame is used for welding purpose.

HARD WATER:

1. Hard water: Water which does not give lather with soap freely is called hard water.
2. Water which gives precipitate with soap is called hard water.
3. Soft water: Water that gives lather with soap is called soft water.
4. Hardness of water is due to presence of bicarbonates, chloride and sulphates of calcium and magnesium.
5. Temporary hardness is due to presence of dissolved bicarbonates of calcium and magnesium.
6. Permanent hardness is due to presence of chlorides and sulphates.
7. Temporary hardness of water can be removed by boiling.
8. By heating with lime also temporary hardness can be removed.(clarks process)
9. In the removal of temporary hardness soluble bicarbonates are converted into insoluble carbonates.
10. Permanent hardness of water can be removed by adding sodium carbonate (soda process)
11. By adding caustic soda both temporary and permanent hardness of water can be removed.
12. Permanent hardness of water can be removed by Calgon process.

13. The substance used in Calgon process is sodium hexa meta phosphate $Na_2[Na_4(PO_3)_6]$.
14. Permanent hardness of water can be removed by Permutit or zeolite method.
15. Sodium aluminum silicate is called sodium zeolite, this is also called sodium aluminum orthosilicate.
16. Artificial zeolite is obtained by the fusion of Quartz (SiO_2), Soda ash (Na_2CO_3) and china clay.
17. The effective method for softening of water used for domestic purposes is zeolite process.
18. When hard water is passed through the permutit the Ca and Mg reacts with the permutit to form complex.
    $Na_2Al_2Si_2O_8\,XH_2O + CaCl_2 \rightarrow CaAl_2Si_2O_8XH_2O + 2NaCl$
19. The exhausted permutit can be regenerated by treating it with brine.
    $CaAl_2Si_2O_8\,XH_2O + NaCl \rightarrow Na_2Al_2Si_2O_8XH_2O + CaCl_2$
20. Permanent hardness of water can be removed by ion-exchange process.
21. Cation exchange resin contains big organic molecules containing acidic groups.
22. Anion exchange resin contains big organic molecules containing basic groups.
23. In cation exchange resin, the cations of water are replaced by H^+ ions.
    $2RCOOH + Ca^{+2} \rightarrow (RCOO)_2\,Ca + 2H^+$
24. In anion exchange resin, the anions of hard water get replaced by OH^- ions.
    $(RNH_3)OH + Cl^- \rightarrow (RNH_3)\,Cl + OH^-$
25. The exhausted cation and anion exchange resins are regenerated by conc. H_2SO_4 and conc. NaOH solutions respectively.
26. Water which contains no H^+ or OH^- ions is called de-mineralized or de-ionized water.
27. Hardness of water is expressed in terms of dissolved salts of calcium and magnesium.
28. If 10^6g of water contains 162g of $Ca(HCO_3)_2$ then the hardness is 100ppm.
29. Hardness of water can be estimated by EDTA (ethylene diamine tetra acetic acid)
30. By using hard water life span of boiler decreases and hard scales are formed in the boilers. Wastage of soap.

HEAVY WATER:

1. Heavy water (D_2O) it is called as deuterium oxide.
2. Heavy water is discovered by Urey.
3. Heavy water is prepared from ordinary water by prolonged electrolysis.
4. 6000 Parts of ordinary water contain 3 part of heavy water.
5. Physical properties of heavy water markedly different from water, but the chemical properties are same due to same electronic configuration
6. Molecular weight of D_2O is 20.
7. Boiling point of heavy water 101.42°C and water is 100°C.
8. Freezing point of heavy water 3.82° is greater than that of ordinary water (0°C)
9. Maximum density of water is at 4°c and heavy water is at 11.6°C
10. Dielectric constant, ionic product and surface tension of heavy water are less than those of ordinary water.
11. The reaction velocity of D_2O is less than that of H_2O.
12. Heavy water is slightly injurious to health.

13. Ionic product of D_2O is less than that of H_2O.
14. Living organisms do not survive in heavy water.
15. Heavy water is used in the preparation of deuterium.
16. Heavy water is used as moderator in nuclear reactor.
17. Heavy water is used as tracer compound.
18. Heavy water is used as pesticide and germicide.

HYDROGEN PEROXIDE (H_2O_2):

1. The compounds in which O-O bond is present are called peroxides.
2. Any metal peroxide treated with ice cold sulphuric acid gives H_2O_2.
3. H_2O_2 can be prepared from Na_2O_2 by the action of ice cold sulphuric acid; this method is called Merck's process.
4. Hydrogen peroxide is manufactured by electrolysis of 50% H_2SO_4 and ammonium sulphate at 0^0c using platinium anode and lead cathode.
5. The product obtained is peroxy di sulphuric acid ($H_2S_2O_8$).
6. The reactions at the electrodes are as follows.
   $2H_2SO_4 \rightarrow 2H^+ + 2HSO_4^-$ (ionization)
   At cathode: (Pb)
   $2H^+ + 2e^- \rightarrow H_2$ (reduction)
   At anode: (Pt)
   $2HSO_4^- \rightarrow H_2S_2O_8 + 2e^-$ (oxidation)
7. In auto oxidation method, the starting substance is 2-ethyl anthraquinone.
8. The reagent used for reducing the 2-ethyl anthraquinone to 2-ethyl anthraquinol is H_2/Pd.
9. On aerial oxidation 2- ethyl anthraquinol gives H_2O_2.
10. The stages present in the concentration of H_2O_2 are freezing of water, evaporation of water, distillation under reduced pressure and crystallization of H_2O_2.
11. In the electrolytic preparation of H_2O_2 hydrogen evolves.
12. H_2O_2 is light blue colored liquid with nitric acid like odour.
13. Density of H_2O_2 is 1.47gr/cc.
14. H_2O_2 forms blisters on the skin.
15. Dilute solution of H_2O_2 is neutral, but the concentrated one is slightly acidic.
16. Equivalent weight of H_2O_2 is 17.
17. H_2O_2 is less stable, and it stability can be increased by adding glycerol or alcohol or acid.
18. Decomposition of H_2O_2 can be increased by MnO_2, Ag, Au, $H_2O_2 \rightarrow H_2O + (O)$
19. H_2O_2 is a weak acid, hence reacts with bases forming corresponding peroxides.
    Ex: $2NaOH + H_2O_2 \rightarrow Na_2O_2 + 2H_2$
20. H_2O_2 oxidises black lead sulphide to white lead sulphate.
    $PbS + H_2O_2 \rightarrow PbSO_4 + H_2O$
21. ous acids are oxidsed to ic acids.
    $HNO_2+H_2O_2 \rightarrow HNO_3+H_2O$
    $H_2SO_3+H_2O_2 \rightarrow H_2SO_4+H_2O$
    $H_3AsO_3+H_2O_2 \rightarrow H_3AsO_4+H_2O$
    $H_3PO_3+H_2O_2 \rightarrow H_3PO_4+H_2O$
22. H_2O_2 liberates iodine from acidified potassium iodide.
    $2KI+H_2SO_4+H_2O_2 \rightarrow K_2SO_4 + I_2 + 2H_2O$

23. Acidified ferrous sulphate is oxidized to ferric sulphate.
    $2FeSO_4 + H_2SO_4 + H_2O_2 \rightarrow Fe_2(SO_4)_3 + 2H_2O$
24. Acidified potassium ferro cyanide is oxidized to potassium ferri cyanide.
    $2K_4\{Fe(CN)_6\} + H_2SO_4 + H_2O_2 \rightarrow 2K_3\{Fe(CN)_6\} + K_2SO_4 + 2H_2O$
25. Acidified potassium dichromate oxidized blue chromium peroxide.(CrO_5)
26. Benzene is oxidized to phenol in presence of ferrous sulphate.
    $C_6H_6 + H_2O_2 \rightarrow C_6H_5OH + H_2O$.
27. H_2O_2 acts as a bleaching agent due to the liberation of nascent oxygen.
28. Bleaching action of H_2O_2 is due to oxidation. It bleaches wool, silk, feathers, Ivory.
29. H_2O_2 bleaches human hair to golden yellow.
30. It reduces silver oxide to silver. $Ag_2O + H_2O_2 \rightarrow 2Ag + H_2O + O_2$
31. Ozone is reduced to oxygen by H_2O_2
32. Alkaline potassium ferricyanide is reduced to potassium ferrocyanide
    $2K_3Fe\{(CN)_6\} + 2KOH + H_2O_2 \rightarrow 2K_4Fe\{(CN)_6\} + 2H_2O$
33. Halogens are reduced to respective halogens.
    $Cl_2 + H_2O_2 \rightarrow 2HCl + O_2$ $Br_2 + H_2O_2 \rightarrow 2HBr + O_2$
34. Acidified potassium permanganate is decolorized by H_2O_2.
    $2KMnO_4 + 3H_2SO_4 + 5H_2O_2 \rightarrow K_2SO_4 + 2MnSO_4 + 8H_2O$
35. Bleaching powder is to calcium chloride
    $CaOCl_2 + H_2O_2 \rightarrow CaCl_2 + H_2O + O_2$
36. Acidified magnous di oxide is reduced to magnous sulphate.
    $MnO_2 + H_2SO_4 + H_2O_2 \rightarrow MnSO_4 + 2H_2O + O_2$
37. Ethylene is converted into ethylene glycol by H_2O_2
    This is addition reaction $C_2H_4 + H_2O_2 \rightarrow C_2H_4(OH)_2$
38. H_2O_2 is used as antiseptic.90% H_2O_2 is used as fuel.
39. H_2O_2 turns starch iodide paper blue.
40. H_2O_2 gives yellow colour with titanium sulphate.
41. H_2O_2 oxidizes acidified $K_2Cr_2O_7$ to blue color CrO_5.
42. H_2O_2 turns acidified TiO_2 solution to orange colored per titanic acid.(H_2TiO_4)
43. Structure of H_2O_2: In H_2O_2 all the four atoms lie in the same plane.
44. It has open book shape.O-H bond angle is 94°.48’.The dihedral angle is 111°.30’.
45. O-O bond distance is 1.48A° and O-H bond length is 0.97A°.
46. In crystalline H_2O_2 the dihedral angle is reduced to 90° due to hydrogen bonding.
47. H_2O_2 forms addition compounds like hyperol with urea $CO(NH_2)_2.H_2O_2$ which gives H_2O_2 with water.
48. Strength of H_2O_2 solution: It is expressed in terms of volume strength, volume percentage, morality & normality. Volume strength: It represents the volume of O_2 liberated on heating H_2O_2 by decomposition.
    Ex: 100vol H_2O_2 means 1 vol (ml or lt) of H_2O_2 on decomposition gives 100vol (ml or lt) of O_2.
49. Merck’s perhydrol contains about 30% (w/v) H_2O_2 corresponding with 100 volume strength. 30% (w/v) H_2O_2 means 30g of H_2O_2 is present in 100 ml solution.
50. The molarity of perhydrol is 8.9M its normality is 17.8N and it contains 300g/lt or 30g/100ml or 0.3g/ml.

HYDRIDES:

1. The binary compounds of hydrogen with another element are known as hydrides.
2. The hydrides are of different type as (1) ionic hydrides (2) metallic hydrides (3) intermediate hydrides (4) molecular hydrides (5) unknown hydrides.
3. The type of hydride depends upon the nature of bond formed between the hydrogen and other element.

Ionic hydrides:

- The hydrides formed by alkali ad alkaline earth metals are known as ionic or salt like or saline hydrides. Ex: NaH, CaH_2, LiH.
- The ionic hydrides are solids with high melting points. They react with acids and water to liberate hydrogen.
- They have high heat of formation and are powerful reducing agents at high temperature. Ex: $CO + NaH \xrightarrow{\Delta} HCOONa + C$
- They are used to remove traces of water from solvents and to prepare other hydride compounds like $LiAlH_4$, $NaBH_4$, etc.

Covalent or molecular hydrides:

- Three are formed by the electromagnetic elements of *p*-block.
- The molecular formulas can be written on MH_n or $MH_{(8-n)}$ where *n* group number. Ex. CH_4, NH_3, H_2O and HF
- They are further classified as
  - (1) Electron – Precise Compounds: All the valence electrons of the central atom are involved in bond formation. Ex: CH_4, C_2H_6.
  - (2) Electron – Deficient Compounds: The central atom as (en number of electron in valence shell than required for normal covalent bond formation. Ex: B_2H_6, $(AlH_3)_n$.
  - (3) Electron – Rich Compounds: The central atom as more the number of electron than required for bond formation. Ex: NH_3, H_2O

Metallic or Interstitial Hydrides:

- There are formed when *d*-block and *f*-block elements react with hydrogen under high pressure. Ex: CrH, CrH_2, ThH_2, ZnH_2.
- They are hard, with metallic luster, good conductors with magnetic properties.
- They are used as catalyst and there densities are less than the parent metal.

Nomenclature of Covalent Hydrides

According to IUPAC the names of covalent hydrides are written by adding suffix ‘ane’ to names of the element.

Formula	IUPAC Name	Common Name
B_2H_6	Diborane	Diborane
CH_4	Methane	Methane (Marsh gas)
NH_3	Azane	Ammonia
PH_3	Phospane	Phospane
AsH_3	Arsane	Arsine
SbH_3	Stibane	Stibine
H_2O	Oxidane	Water
H_2S	Sulfane	Hydrogen sulfide
HF	Hydrogen flouride	Hydrogen fluoride
HCl	Hydrogen chloride	Hydrogen chloride

HYDRIDES

SYNOPSIS:-

1. The Binary compounds of hydrogen with other elements are called hydrides.
2. Based on stability, hydrides are divided into five types.
   1. **Ionic hydrides:** Except radio active metals, Be and Mg the remaining S-block metallic hydrides are called Ionic hydrides.
   2. **Intermediate hydride:** Be, Mg, Cu, Zn and Al hydrides are called Intermediate hydrides.
   3. **Metallic hydrides:-** IIIB (Lanthanide and actinides), IVB, VB and Cr hydrides are called metallic hydrides.
   4. **Unknown hydrides:-** Pd, Bi and Po metal hydrides are called unknown hydrides.
   5. **Molecular hydrides**:- Except zero group elements, Al, Bi, & Po the remaining p-block.
3. In the periodic table the region of groups 7 to 9 is referred to as the hydride gap because metal in these group do not form metallic hydrides.

The types of hydride formed depends on the electro negativity difference between hydrogen and other element.

The hydrides are mainly classified into three types.

1) Saline (or) Ionic (or) salt like hydrides
2) Molecular (or) covalent hydrides
3) Metallic (or) Interstitial hydrides

Ionic (or) Saline (or) salt like hydrides:-

These are the hydrides formed between hydrogen and metals of low electro negativity (or) electropositive metals of group IA and IIA (Except Be, Mg & Radioactive metals)

- S-block metals $+ H_2 \xrightarrow{300^{o}C} Ionic\, hydride$
- Examples:- NaH, KH, CaH_2
- These are non-volatile crystalline solids, having white (or) grayish colour.
- They have high melting point & boiling point. Thermal stability of alkali metallic hydrides decreases from LiH to CsH.
- They have high density and high heat of formation Stoichiometric Ionic hydrides are insoluble in non-aqeous solvents.
- All ionic hydrides react with acids and water vigorously and liberates H_2 gas.
- $Li\,H + H_2O \rightarrow Li\,OH + H_2\uparrow$
- CaH_2 is known as hydrolith.
- $CaH_2 + 2H_2O \rightarrow Ca(OH)_2 + 2H_2\uparrow$
- Except Li H remaining Ionic hydrides decomposition at 400 – 500°C produces metals and H_2 gas.
- Electrolysis of molten hydrides produces H_2 gas at anode.

  $MH \rightleftarrows M + H^-$

  Anode reaction, Cathode reaction

  $2H^- \rightarrow H_2 + 2e^-$ $M^+ + e^- \rightarrow M$

  This oxidation reaction confirms that presence of H- ions.

The density of these hydrides is greater than that of metal from which they are formed.

At high temp ionic hydrides are powerful reducing agents due to the presence at nascent hydrogen atoms.

$2CO + NaH \xrightarrow{\Delta} Si\ H_4 + 4\ NaCl$

Ionic hydrides are fused with Intermediate hydrides (or) molecular hydrides and produces complex hydrides.

$Li\ H + Al\ H_3 \rightarrow Li[Al\ H_4]$

$2NaH + B2\ H6 \rightarrow 2Na\ [BH_4]$

Complex hydrides are powerful reducing agents than their constituent compounds.

They reacts vigorously with water and other protic solvents such as long alcohol and ammonia liberates H^2 gas.

$NaH + C_2H_5\ OH \rightarrow Na\ OC_2H_5 + H_2$

$NaH + NH_3 \rightarrow Na\ NH_2 + H_2$

Uses: They are used in the laboratory to remove traces of water from solvents and inert gases such as N_2 and Ar.

2. In the preparation of complex hydrides.

Metallic (or) Interstitial Hydrides:

- Metallic or interstitial hydrides are formed by and f-block elements.
- Metallic hydrides are formed by the presence of hydrogen atoms in the position of the metal crystals.
- Examples Ti $H_{1.7}$, Zr $H_{2.5}$
- Metallic hydrides are non-stiochiometric compounds and show electric conduction.
- The properties of metallic hydrides are less than those of the parent metals.
- The interstitial hydrides release hydrogen when heated and this property is used in the purification and for storing H_2.

Molecular (or) Covalent Hydrides:

- **Covalent hydrides are formed by p-block elements.**
- **Covalent hydrides are formed usually consist of individual, discrete covalent molecules.**
- **The molecular formula of covalent hydrides can be written as MH_n or $MH_{(8-n)}$ where n = group number of the element is short form period table.**

S.No.	Formula IUPAC	Name of the hydride Common	The hydride
1.	**$B_2\ H_6$**	**Diborane (6)**	**Diborane**
2.	**CH_4**	**Methane**	**Methane**
3.	**NH_3**	**Azane**	**Ammonia**
4.	**PH_3**	**Phosphane**	**Phosphine**
5.	**AsH_3**	**Arsane**	**Arsine**
6.	**SbH_3**	**Stibane**	**Stibine**
7.	**H_2O**	**Oxidane**	**Water**
8.	**H_2S**	**Sulfane**	**Hydrogen sulfide**
9.	**HF**	**Hydrogen**	**Hydrogen fluoride**
10.	**HCl**	**Hydrogen**	**Hydrogen chloride**

- Covalent hydrides are named from the name of element and the suffix-ane.
  Ex: Phosphane – PH_3 : Oxidane H_2O an azane – NH_3
- Molecular hydrides are three types i) electron deficient ii) electron – precise and iii) electron – rich.
- Electron – deficient molecular hydrides have less number of electrons for writing and conventions .
  Lewis structure Ex: B_2H_6
- Electron – precise molecular hydrides are forms by the elements of IV or 14 group elements in which all the valence electrons are involved in bond to motion
- Electron rich molecular hydrides contain more valency electrons than required for bond formatic i.e., they contain lone pairs of electrons Ex: NH_3 H_2O.
- Between the molecules of covalent hydrides were vander waal's attractive forces exists, hence they are volatile with low boiling points.
- The presence of lone pair of electrons on high electronegative atoms like F,O,N in electrons rich molecular hydride result in hydrogen bond formation, which leads to high boiling points, high primitives and aggregation.
  Be, Mg, and Al forms polymeric hydrides.
  Examples: *$(BeH_2)x$, $(Mg\ H_2)x$ $(Al\ H_3)$*

Question Bank:

1. The number of protons present in Deuterium are
(1) 1 (2) 2 (3) 3 (4) 4

2. The electronic configuration of deuterium atom is
(1) $1s^2\ 2s^0$ (2) $1s^1\ 2s^0$
(3) $1s^2\ 2s^2$ (4) $1s^1\ 2s^1$

3. The H-O-H bond angle in water molecule is
(1) $103^\circ\ 50^1$ (2) 105°
(3) $104^\circ\ 30^1$ (4) $105^\circ\ 28^1$

4. The percentage of Deuterium in Heavy water
(1) 22 (2) 11.11
(3) 4 (4) 20

5. The number of electrons, protons and neutrons in tritium are in the ratio of
(1) 1 : 1 : 1 (2) 1 : 1 : 2
(3) 2 : 1 : 1 (4) 1 : 2 : 1

6. Permanent hardness of water is due to
(1) Na, K (2) Mg, K
(3) Ca, Mg (4) Ca

7. Which of the following salt is responsible for temporary hardness of waters
(1) calcium bicarbonate
(2) calcium chloride
(3) calcium nitrate
(4) calcium sulphate

8. which of the following atoms are present in heavy water
(1) $_1H^2\,_8O^{16}$ (2) $_1H^1\,_8O^{18}$
(3) $_1H^3\,_8O^{17}$ (4) $_1H^1\,_8O^{16}$

9. Which of the following is a peroxide
(1) Cu_2O_2 (2) MnO_2
(3) PbO_2 (4) Na_2O_2

10. O-O- bond is present in
(1) peroxide (2) dioxide
(3) ether (4) carboxylic acid

11. Potassium ferrocyanide is converted to ferricyanide by

H_2O_2, then oxidation state of Fe changes from _____ to _____

(1) $Fe^{+2} \rightarrow Fe^{+3}$
(2) $Fe^{0} \rightarrow Fe^{+2}$
(3) $Fe^{+3} \rightarrow Fe^{+2}$
(4) $Fe^{+2} \rightarrow Fe^{+}$

12. The concentration of 10 volume H_2O_2 in weight present is

(1) 3% (2) 1%
(3) 10% (4) 30%

13. H_2O_2 reacts with Ag_2O and gives

(1) Ag_2O_2 (2) Ag
(3) AgOH (4) Ag_2CO_3

14. Hydrogen cannot be obtained by

(1) Zn and dil H_2SO_4
(2) Mg and conc.HNO_3
(3) Mg and H_2SO_4
(4) Mg and H_2O

15. Calculate the normality of 10 volume H_2O_2

(1) 1.7N (2) 12N
(3) 30.3N (4) 0.303 N

16. H_2O_2 oxidises H_2S to

(1) S (2) H_2SO_4
(3) $H_2S_2O_8$ (4) H_2SO_5

17. The weight of one mole of heavy water

(1) 18 gm (2) 20 gm
(3) 22 gm (4) 6.03×10^{23}gm

18. In the presence of ferrous sulphate, H_2O_2 oxidises Benzene to

(1) Benzaldehyde
(2) Benzoic acid
(3) Aniline
(4) Phenol

19. The formula of the compound obtained when H_2O_2 reacts with PbO_2

(1) $PbSO_4$ (2) PbO
(3) PbO_2 (4) Pb

20. H_2O_2 converts black PbS into white coloured

(1) PbO (2) $PbSO_4$
(3) PbO_2 (4) Pb

21. If calcium chloride is present in a dissolved state in water, the water can be called as

(1) soft water
(2) hard water
(3) impure water
(4) chlorinated water.

22. The hardness of water due to Magnesium bicarbonate is

(1) temporary hardness
(2) permanent hardness
(3) soft water
(4) none

23. When water is purified by ion exchange method, the cation exchange resin exchanged the ion from water is

(1) H^{+} (2) OH^{-}
(3) Metal ions (4) Cl^{-}

24. H_2O_2 forms two types of salts, peroxide and

(1) Superoxides
(2) Oxides
(3) Hydroperoxides
(4) All the above

25. H_2O_2 converts most silver oxide to

(1) silver peroxide
(2) silver hydroxide
(3) silver
(4) all the above

26. The concentration of H_2O_2 can be increased by
(1) distillation under reduced pressure
(2) steam distillation
(3) fractional distillation
(4) crystallization

27. The weight of H_2O_2 in 10 ml of 20 volume of H_2O_2 solution is
(1) 0.303 gm (2) 0.6 gm
(3) 6 gm (4) 60 gm

28. The magnetic property of H_2O_2 is
(1) diamagnetic
(2) paramagnetic
(3) ferromagnetic
(4) none

29. The organic compound used in the preparation of H_2O_2 by auto oxidation method is
(1) anthracene
(2) anthraquinone
(3) quinine
(4) any of the above

30. Bleaching property of H_2O_2 is due to
(1) Oxidising property
(2) Reducing property
(3) Acidic property
(4) Dehydration

31. The reaction of heavy water with $AlCl_3$ is called
(1) Hydrolysis (2) Deuterolysis
(3) Ozonolysis (4) Hydration

32. If the normality of H_2O_2 is 1.78N then concentration in volumes is
(1) 10 volumes (2) 100 volumes
(3) 1 volume (4) 177 volumes

33. Water can be detected by
(1) Adding phenophthalein solution
(2) Blue litmus
(3) Red litmus
(4) Anhydrous $CuSO_4$ salt.

34. The bottle of H_2O_2 marked as 20 volumes indicates that
(1) 20 ml of H_2O_2 on decomposition gives are lit of O_2
(2) one vol of O_2 is obtained from 20 vol of H_2O_2
(3) one lit of H_2O_2 on decomposition gives 20 lit of O_2
(4) one ml of H_2O_2 solution at STP gives 20 ml of O_2

35. Which of the following does not cause hardness to water?
(1) NaCl (2) $CaCl_2$
(3) $MgSO_4$ (4) $Ca(HCO_3)_2$

36. Which does not represent soap?
(1) Sodium stearate
(2) Sodium lauryl sulphate
(3) Sodium palmiate
(4) Sodium oleate

37. Which is insoluble in water
(1) $Ca(HCO_3)_2$
(2) $Mg(HCO_3)_2$
(3) $MgCO_3$
(4) NaCl

38. When the water contained temporary hardness is boiled, it releases a gas *X*. Then the rate of diffusion of *X* is same as that of
(1) NO_2 (2) N_2O
(3) N_2 (4) C_2H_6

39. When Zeolite is added to hard water the sodium ions are exchanged with
(1) H^+ ions (2) Ca^{2+} ions
(3) SO_4^{2-} ions (4) OH^- ions

40. Hard water becomes free from ions when passed through ion

exchange resin containing RCOOH groups

(1) Cl^- (2) $SO_4^{2\pm}$
(3) H_3O^+ (4) Ca^{2+}

41. Gan permutit method is also called
(1) Calgon method
(2) ion – exchange method
(3) base – excahnge method
(4) none of the above

42. Assertion: The reactivity of Heavy water is less than that of H_2O.
Reason: The –O bond is stronger than H –O bond.
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

43. Assertion: Water containing higher concentration of D_2O is not poisonous
Reason: Seeds do not germinate and small fish die in heavy water
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

44. Assertion: On electrolysis, A solution of Heavy water containing Na_2CO_3 liberates deuterium at cathode and oxygen at anode
Reason: Na_2CO_3 enhances the conductivity of Heavy water.
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

45. Assertion: Heavy water also forms salt deuterates like ordinary water.
Reason: Minute quantities of Heavy water are found in rain water
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

46. Assertion: O_2 gas is liberated in all the oxidizing properties of H_2O_2.
Reason: Delicate material like silk, wool etc. are bleached by H_2O_2.
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

47. Assertion: MnO_2 cause rapid decomposition of H_2O_2.
Reason: MnO_2 acts as catalytic poison.
(1) Both **A** and **R** are true and R is the correct explanation of **A**

(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

48. Assertion: With a solution of TiO_2 in conc. H_2SO_4, H_2O_2 gives orange colour.
Reason: Pertitanic acid is a orange yellow compound.
(1) Both **A** and **R** are true and R is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false.
(4) **A** is false but **R** is true.

49. Match the following.

List-I	**List-II** ***Melting Point***
A) H_2O_2	1) –254.5°C
B) H_2	2) 3.82°C
C) D_2O	3) –259.2°C
D) D_2	4) –0.4°C
	5) 100°C

The correct matching List-I and List-II is

	A	**B**	**C**	**D**
(1)	4	3	2	1
(2)	1	2	3	4
(3)	5	3	4	1
(4)	2	5	1	4

50. Match the following

List-I ***Starting compound***	**List-II** ***Process of preparation of H_2O_2***
A) $BaO_2+8H_2O_2$	1) electrolysis
B) 50% H_2SO_4 (Cold)	2) Auto oxidation
C) 2-ethyl anthra quinol	3) laboratory method
D) 2- ethyl anthraquinone	

The correct matching from List-I and List-II is

	1	**2**	**3**
(1)	D	A	C
(2)	B	D	A
(3)	A	C	D
(4)	A	D	C

KEY

1) **1** 2) **2** 3) **3** 4) **4** 5) **2** 6) **3** 7) **1** 8) **1** 9) **4** 10) **1**
11) **1** 12) **1** 13) **2** 14) **2** 15) **1** 16) **1** 17) **2** 18) **4** 19) **2** 20) **2**
21) **2** 22) **1** 23) **3** 24) **3** 25) **3** 26) **1** 27) **2** 28) **1** 29) **2** 30) **1**
31) **2** 32) **1** 33) **4** 34) **4** 35) **1** 36) **2** 37) **3** 38) **2** 39) **2** 40) **4**
41) **3** 42) **1** 43) **4** 44) **1** 45) **2** 46) **4** 47) **3** 48) **1** 49) **1** 50) **2**

Objective Type Questions

1. $Ti\ H_{1.73}$ is an example of
(1) ionic hydride (2) covalent hydride
(3) metallic hydride (4) polymetric hydride

2. Hydrogen does not combine with
(1) He (2) Na (3) Sb (4) Bi

3. Hydrogen directly combines with
(1) Ca (2) Mg (3) Al (4) Cu

4. Which hydride is an ionic hydride?
(1) NH_3 (2) H_2S (3) $Ti\ H_{1.73}$ (4) NaH

5. $Si\ H_4$ is an example of which of the following of hydrides.
(1) ionic (2) interstitial (3) metallic (4) covalent

6. Which among the following is a hydride
(1) Chloride of lime (2) nitrolim (3) hydrolith (4) minium

7. The group of elements readily to produce sal compounds.
(1) IA (2) II A (3) III A (4) 1 & 2

8. In the following compounds H is covalent bond in case of
(1) $Ba\ H_2$ (2) CaH_2 (3) (4)

KEY

1) **3** 2) * 3) * 4) * 5) * 6) * 7) * 8) **4**

9. ALKALI AND ALKALINE EARTH METALS

Synopsis:

1. Elements – Lithium (Li), Sodium (Na), Potassium (K), Rubedium (Rb), Cesium (Cs) Francium (Fr). Francium is radioactive.
2. Alkali metals are not found in Free State. They form water soluble hydroxides which are strongly alkaline hence known as alkali metals.
3. Alkali metals belong to s- block. There ground state configuration is ns^1.
4. Physical state- Soft, silvery white.
5. Metallic bond – Strength of metallic bond is weak, possess low cohesive energy, can be cut easily. Strength of metallic bond decreases from Li to Cs.
6. Stability:- Alkali metals are highly reactive, they are stored under kerosene or petrol.
7. All metals have low I.P. the metallic nature increases from Li to Cs, because I.P. decreases from Li to Cs. All elements are strongly electropositive
8. Atomic radii:- Alkali metals have largest atomic size in the period. Atomic and ionic radii increase from Li to Cs. Addition of extra shell & screening effect which also increases atomic volume.
9. Density : They possess low density because of large atomic size and weak metallic bonding density increases from Li to Cs but density of Na>K. Li is the lightest known metallic element ($d = 0.53$ g/m).
10. M.Pt. and B.Pt : It decreases down the group from Li to Cs.
11. Oxidation state:- They easily lose outermost electron to form uni postive M^+ ion . O.S is +1 all alkali metal ions Li^+ Na^+ K^+ Rb^+ Cs^+ are diamagnetic and colour less.
12. Electro negativity: E.N. values are low. E.N. decreases from Li to Cs.
13. Photo electric effect: Because of electro positive or metallic character they emit electrons when exposed to light. This phenomenon is called photo electric effect. K and Cs are used in photo electric cells.
14. Hydration of ions:- Solubility of alkali metal ions is due to hydration of cations by water molecule smaller the size of cation, greater the degree of hydration. Li^+ is more hydrated than Cs^+ because of this Li^+ is best conductor in aqueous medium.
15. Flame coloration:- Alkali metals and salts introduced in flame of bunsen burner give flame colour, due to electronic transition in outer energy levels.

Alkali metal	Colour of flame
Li	Crimson red
Na	Golden yellow
K	Pale pink
Rb	red violet
Cs	Blue

16. General chemical properties: Chemical reactivity Increases from Li to Cs because of decrease of cohesive force and decrease in strength of metallic bond.
17. Oxides : On exposed to air, surface of alkali metals get tarnished due to formation of oxide film with air metals forms monoxides, hydroxides (In moisture), Carbonates (with CO_2).

18. Alkali metals catch fire in air and burn to produce monoxide (M_2O). Therefore they are preserved in kerosene. Li burns in air to produce lithium nitride also.
19. Alkali metals can form monoxide M_2O, peroxide M_2O_2 and super oxide MO_2. Li does not form peroxide and super oxide due to strong polarizing power of its ion. Li^+ ion is strong enough to restrict the oxidation of an oxide ion O^{-2} to O^{-}_2 peroxide ion.
20. Sodium does not form super oxide, because Na^+ prevents oxidation of O^{-2} to O_2^{-}
21. Monoxides have O^{-2} ion, peroxides O_2^{-2} ion and superoxides ion O_2^{-}. O.N of oxygen in monoxide =-2 peroxide = 1, super oxide - ½ .
22. Peroxides are diamagnetic, Na_2O_2 is called oxone – used for bleaching wood pulp, paper, fabrics. It is a strong oxidizing agents.
23. Superoxides: Contains odd electron, they are paramagnetic and coloured. Super oxide reacts with water liberating H_2O_2 and O_2.
24. CsO_2 and KO_2 are orange is coloured super oxides are strong oxidizing agents.
25. Action of water: They react readily and vigorously to form hydroxides with the liberation of Hydrogen. Reactivity towards water increases due to increases in electropositive nature.
26. Hydroxides are basic in nature crystalline solids, soluble in water and alchohol.
    $M_2O+H_2 \rightarrow MOH$; $M+ H2O \rightarrow MOH + \frac{1}{2} H_2$
    Basic strength o hydroxides increases from lithium to cesium.
27. With liqud ammonia: - All alkali metals are soluble in liquid ammonia to form blue colored solution is stable and conducts electricity due to solvated electrons.
28. Carbonates and bicarbonates: All alkali metals form carbonates (M_2CO_3) and bicarbonates ($MHCO_3$) carbonates are weak acids. Alkali metals are only elements which form most stable solid bicarbonates.
29. Nitrates: All alkali metals form nitrates (MNO_3) nitrates are highly soluble in water, give nitrites on strong heating.
30. Amalgams: Alkali metals form amalgams with mercury. Eg: NaHg, Na_2Hg, Na_3Hg.
31. Abnormal behavior of lithium: Lithium differs form other alkali metals It is due to small size, high E.N. and I.P. than other alkali metals, comparatively hard, high polarizing effect, absence of vacant *d*-orbital. Li-resembles Mg due to diagonal relationship, because polarizing power of Li and Mg are same.

SODIUM:

1. Sodium is the most abundant alkali metal in earths crust. First isolated by Davy by electrolysis of fused caustic soda (NaOH) do not occur in free state occurs as rock salt Nacl, found in sea water and lakes.
    Sodium nitrate ($NaNO_3$) — Chile salt Petre
    Na_2CO_3 $10H_2O$ – Washing soda.
    $Na_2SO4.10H_2O$ – Glaubers salt.
    $Na_2B_4O_7$ 10 H_2O – Borax
    Na_3AlF_6 – Cryolite
2. Extraction of Sodium by Castner's process:
    Electrolyte is Molten NaOH
    Cathode – iron rod – product – sodium
    Anode – Nickle or iron – oxygen gas

3. Properties of sodium:
   - (i) Soft silvery white metal-lighter than water, highly reactive.
   - (ii) It is highly electropositive liberates Hydrogen with acids.
   - (iii) Reacts with all non-metals except (C & N).
   - (iv) Reacts with H_2 to form NaH with Cl_2 to form NaCl with S gives Na_2S. With P gives Na_3P.
   - (v) It is a good reducing agent. It reduces SiO_2 to Si, Al_2O_3 to Al, CO_2 to Na_2CO_3.
4. Blue coloured solution of Na / Liq NH_3 is used as reducing agent When heated with mercury, forms sodium amalgam of varying composition as NaHg, Na_2Hg & Na_3Hg.
5. Uses of sodium
   - (i) In manufacture of chemicals – Na, NaOH, NaCN etc.
   - (ii) In industries for production of dyes, drugs, perfumes, artificial rubber and as a catalyst.
   - (iii) In Lassaign's test for detection of N, S, Cl inorganic compounds.
   - (iv) In Wurtz reaction for preparation of alkanes.
   - (v) In preparation TEL. (anti-knocking agent)
6. Sodium hydroxide (NaOH) caustic soda:
   - (i) Deliquescent white solid. Highly soluble in water with evolution of heat, due to formation of hydrates NaOH nH_2O.
   - (ii) Caustic soda & caustic potash have causticising property i.e. can breakdown the protein of skin and flesh into pasty mass.
7. Manufacture of NaOH:
   - (i) Causticising Process – Action of sodium carbonate and milk of lime gives NaOH.
     $NaCO_3 + Ca(OH)_2 \rightarrow 2NaOH + CaCO_3$
   - (ii) Electrolytic process:- Nelson cell method or Porous diaphragm process:
   - (iii) Electrolyte – Brine solution (NaCl)
   - (iv) Cathode - perforated steel vessel - product NaOH.
   - (v) Anode — graphite rod — product Cl_2.
   - (vi) NaOH is formed in remaining solution.
   - (vii) Castner – Kellener method (or) Mercury cathode cell method:
8. Outer compartments: brine solution, anode graphite rod – Cl_2, cathode Hg – Na (s).
9. Middle compartment – dilute sol. of NaOH. Anode is Hg – H_2 ; Cathode is iron rods – NaOH.
10. Properties of NaOH:
    - (i) Deliquescent solid, soapy to touch, corrosive to skin.
    - (ii) It neutralizes all acids and acidic oxides forming salts because it is a strong alkali.
    - (iii) Al, Zn, Pb, Sn, Cr, Sb are metals which dissolve in NaOH solution liberating H_2.
    - (iv) Aqueous solution of NaOH & KOH absorbs CO_2 from atmosphere.
    - (v) Alcoholic KOH is used as dehyro halogenating agent.

11. Uses of NaOH:
    - (i) in preparation of soda lime, manufacture of soap, dyes, paper, rubber and plastics.
    - (ii) in metellurgy and petroleum industry, mercerizing cotton.
    - (iii) In preparation of bleaching agent NaOCl.
12. Sodium carbonate (Na_2CO_3)
    Anhydrous Na_2CO_3 – soda ash
    Crystal carbonate – Na_2CO_3, H_2O
    Decahydrate – Na_2CO_3 $1OH_2O$ (washing soda)
    Heptahydrate – Na_2CO_3 $7H_2O$
13. Preparation of Na_2CO_3 by (a) leblanc's process (b) Solvay process (c) electrolytic process, In leblane process raw materials used are Nacl, coke, lime stone, H_2SO_4, ammonia solvay process is based on low solubility of $NaHCO_3$. the raw materials used are $CaCO_3$, NH_3 and Nacl.
14. Properties Na_2CO_3 & $NaHCO_3$:
    - (i) Both Na_2CO_3 and $NaHCO_3$ liberate CO_2 gas with dil mineral acids and carboxylic acids.
    - (ii) $NaHCO_3$ is an acidic slat. But aqueous solution of this is alkaline in nature due to anionic hydrolysis.
    - (iii) $NaHCO_3 + H_2O \rightarrow Na^+ + OH^- + H_2CO_3$
    - (iv) Na_2CO_3 reacts with SO_2 to give Na_2SO_3, SiO_2 to give Na_2SiO_3, $MgCl_2$ to give $MgCO_3$
    - (v) It is used in manufacture of glass, water glass and for softening water in paper and dye industries.
    - (vi) Sodium bicarbonate $NaHCO_3$ (Baking soda): Less soluble in water than Na_2CO_3.
    - (vii) All baking powders contain $NaHCO_3$ and acid salt such as potassium hydrogen tartarate or calcium dihydrogen phosphate.
    - (viii) Equimolar mixture of Na_2CO_3 and K_2CO_3 is called fusion mixture.
    - (ix) $NaHCO_3$ is used to make breads, cakes light an fluffy because when it is heated CO_2 is evolved.

ALKALI EARTH METALS

Synopsis:

1. Be, Mg, Ca, Sr, Ra are called alkaline earth metals, and also called IIA group elements
2. Radium is a radioactive element. Alkaline earth metals are malleable, ductile, silvery white lustrous metals.
3. Outermost electronic configuration of these elements is ns^2.
4. All these element show +2 oxidation state in their compounds.
5. Flame coloration Be, and Mg will not give flame colorations Ca-Brick red; Sr-crimson; Ba-apple green.
6. Electropositive of this element is very high. They are less electro positive than alkali metals.
7. IIA group oxides are basic. Hence they are called alkaline earth metals.
8. Among the oxides of IIA group elements BeO is amphoteric and the remaining are basic oxides.

9. Among the halides $Becl_2$ is covalent.
10. Hydration energy of IIA group element decreases from Be to Ba.
11. Among the halides $BeCl_2$ is covalent MgH_2 is covalent to some extent, and the remaining are ionic hydrides.
12. Alkaline earth metals are not as soft as alkali metals.
13. The stability of carbonates, sulphates, of these elements increases down the group.
14. The solubility's of carbonates, sulphates, halides and nitrates decreases down the group.
15. Electro negativity and ionization potential decrease from Ba to Ra.
16. Alkaline earth metals are in most instances less soluble in water than the corresponding alkali metals this due to less size and more charge of alkaline earth metals.
17. Lattice energy of alkaline earth metal compounds are larger than that of the alkali metals which lead to decreases stability.
18. Be does not react with water even at 1000°c and Mg reacts with hot water.
19. Reactivity of IIA group elements increases down the group.
20. Magnesium: It does not occur in the Free State. Important compounds of magnesium are

Magnesite	$MgCO_3$
Carnallite	$KCl, MgCl_2 6H_2O$
Bruicte	$Mg(OH)_2$
Dolomite	$MgCO_3 CaCO_3$
Epsom salt	$MgSO_4 7H_2O$
Asbestos	$CaMg_3 (SiO_3)_4$

21. Green coloured matter in plants is called chlorophyll contain magnesium.
22. Sea water contains magnesium in the form of $MgCl_2$ in the extraction of Mg by thermal reduction of MgO, the substances used for the reduction are ferrosilicon or calcium carbide.
23. Magnesium is extracted by electrolysis.
24. In electrolytic preparation of Magnesium iron tank acts as cathode and graphite rod act as anode. Electrolysis is carried at 700^0c
25. Magnesium chloride is taken as electrolyte in the electrolytic cell.
26. To lower the melting point of $MgCl_2$ small amount of NaCl is added.
27. Magnesium burns in oxygen with dazzling white light.
    $2Mg + O_2 \rightarrow 2MgO$
28. Magnesium combines with nitrogen to form magnesium nitride.
    $3Mg + N_2 \rightarrow Mg_3N_2$
29. Mg reacts with dilute acids liberating hydrogen.
    $Mg + 2HCl \rightarrow MgCl_2 + H_2$
    $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$
30. With concentrated H_2SO_4 magnesium liberates SO_2.
31. With concentrated HNO_3 it gives NH_4NO_3.
32. Magnesium has no action with alkalis.
33. Magnesium acts as reducing agent, it reduces CO_2 to carbon.
    $CO_2 + Mg \rightarrow 2MgO + C$
34. Magnesium continues to burn even in the CO_2 to carbon.

35. Magnesium is used in the preparation of Grignard's reagent.(RMgX).

$CH_3I + Mg \xrightarrow{\text{Dry ether}} CH_3MgI$

36. Magnesium powder along with potassium chlorate or Barium peroxide is used in flash light photography.
37. Magnesium is used in optical mirrors.
38. Calcium: The ores of Calcium are $CaCO_3$ – Lime stone; CaF_2 - Flourspar,

$CaF_2\ Ca_3(PO_4)_2$ – Fluorapatite; $CaSO_4\ 1/2\ H_2O$ - Plaster of paris

$CaSO_4\ 2H_2O$ - Gypsum

39. The ores of Barium are Barytes($BaSO_4$), Whtherite($BaCO_3$)
40. Calcium sulphate ($CaSO_4$); dihydrated calcium sulphate is called gypsum ($CaSO_4$, $2H_2O$) and semi hydrated calcium sulphate is called plaster of paris. ($CaSO_4.1/2\ H_2O$)
41. CaO or $Ca(OH)_2$ or $CaCO_3$ reacts with concentrated H_2SO_4 to give $CaSO_4$.

$CaO+H_2SO_4 \rightarrow CaSO_4 +H_2O$

$Ca(OH)_2 +H_2SO_4 \rightarrow CaSO_4+2H_2O$

$CaCO_3+H_2SO_4 \rightarrow CaSO_4+ CO_2 +H_2O$

42. Gypsum is insoluble in water but soluble in ammonium sulphate and its solubility in water decreases with increase of temperature.
43. When Gypsum is heated to 120°c it loses water of crystallization and converted into plaster of paris.

$2(CaSO_42H_2O) \rightarrow (CaSO_4)_2\ H_2O + 3H_2O$

44. Completely dehydrated gypsum is called dead burnt obtained at 200°c.
45. Gypsum is used in chalk pencils, baking powder, tiles, plasters and fertilizers.
46. Gypsum also is used in port land cement to prevent rapid setting.
47. Plaster of paris ($CaSO_41/2\ H_2O$) or $(CaSO_4)_2H_2O$.
48.
    (i) Rate of setting of plaster of paris is increased by common salt and decreased by borax.
    (ii) Plaster of Paris on setting gives a hard mass of gypsum.
    (iii) The setting involves the two stages 1. Setting stage and 2. Hardening stage.
    (iv) in the setting of plaster of paris, it is converted of Orthombic dehydrate which hardenes to monoclinic gypsum.
49. Quick lime reacts with water to give slaked lime. $CaO + H_2O \rightarrow Ca(OH)_2$
50. When quick lime is heated with coke at high temperature it forms calcium carbide. $CaO+3C \rightarrow CaC_2 + CO$
51. It is used as drying agent and flux in metallurgy.
52. Mortar: Mortar is the mixture of slaked like (1 part) and sand (3 parts).
53. Mortar is used in the construction of buildings.
54. Mortar mixed with cement is known as cement mortar which is harder than lime mortar.
55. Lime stone and clay containing 10% aluminium silicate when heated together gives “hydraulic mortar“.
56. Hydraulic mortar is used for bleaching purposes and as antiseptic.
57. A setting of mortar is facilitated of $AlCl_3$.
58. Cement is rich with lime and silica

59. The raw materials for manufacture of cemtn are limestoe and caly.
60. Cemnt clinker is a mixture of calcium silicates and calcium aluminates.
61. Gypsum (2 to 3%) is added to cement to slow down the process of setting of cement.

Anamolous Properties of first elements

- Li is hard metal while other alkali metals are soft.
- It has high melting point & boiling point.
- Li only unites with N_2 forming nitrides. $6Li+N_2 \rightarrow 2Li_3N$
  - Li only forms carbide.
  - The solubility of Lithium compounds are less composed to other compounds.
  - Li OH is weak alkali; least basic.
    It is different due to its small size, forms greater covalent bonds and also form co-ordination compounds.

DIAGONAL RELATIONSHIP OF LI:

- Li shows diagonal relation ship with Mg due to similar polarizing powers, electronegativites.
- Li reactly slowly with water. Mg decomposes in hot water.

  $2Li+2H_2o \rightarrow 2LiOH+H_2$

  $Mg+2H_20 \rightarrow Mg(OH)_2+H_2$
- Li & Mg combines with N_2 forming nitride.
- Li Cl_2 & Mg Cl_2 are deliquescent but Li Cl undergoes hydrolysis to smaller extent like Mg Cl_2.
- Li & Mg halides are covalent and are soluble in organic solvents.
- Both Li^+ & Mg^{+2} are highly hydrated.
- The carbonates, phosphates & fluorides of both Li & Mg are sparingly soluble in water.
- Li alkyls are chemically similar to Grignard's reagent in organic synthesis.

Use of Alkali Metals:

Potassium: An alloy of Na & K is used in high temperature thermometers.

- K is used in photo electric cells.
- KOH is used in soft soaps.
- Used as an electrolyte in storage batteries.

Preparation and Properties of NaCl

Common salt or rock salt is found is sea water and is obtained by evaporation.

Properties of NaCl

- Ordinary NaCl is hydrogroscopie due to presence of calcium & magnesium chlorides.
- It is solubility changes with charge in temperature.
- It has FCC lattice strong.

Uses of NaCl

- It is essential constituent of diet.

- It is startial material for manufacture of Na in Down's process, NaoH in Castner – Kelner process & washing soda in solvay process.
- In the preparation of freezing mixture with ice.
- As a preservative for meat & fish.

Biological importance of Na & K

- There are 27 elements of metals K non-metals in the living system of these Na, K, Mg and Ca are required in major quantities.
- Ions maintain asmotic pressure of cell.
- The ion transport activity is known as sodium pump. But K^+ are not pumped out. The energy required for pumping out Na^+ or taking in K^+ is provided by hydrolysis of ATP into ADP.
- Na+ ions is associated with movement of glucose into cells.

The potassium ions for metabolism of glucose inside the cell.

Anamolous Properties of Beryllium:

- It differs from other elements due to its small size and high electronegativity.
- Be compounds are covalent due to its high polarizing power.
- Be is not affected by air and does not decompose water at ordinary temperature.
- Be is amphoteric metal. It dissolves in alkali.
- $BeSo_4$ is only soluble in water.
- Be & its salts does not respond to flame test.
- Be only forms complexes.
- Be has maximum covalency 4.

Diagonal relationship of Be:

Beryllium shows diagonal relationship to Aluminium.

- Be KAI have some B, N polarizing power.
- They form covalent compounds and are soluble in organic solvents.
- Be, Al compounds undergo hydrolysis.
- Be, Al forms complexes.
- Be, All are passive to Conc. HNo_3.
- Be & Al are atmospheric & dissolve in alkalies and from beryllates, aluminates.
- Bc & A1 form carbides.

Biological importance of Mg and Ca

Mg^{+2}

1) Mg^{+2} ions are concentrated in animal cells.
2) Enzymes like "Phosphohydrolases and phosphotrans ferals contain Mg^{+2} ions.
3) Mg^{+2} is a constituent of Chlorophyll.

Ca^{+2}

Ca^{+2} is present in bones & teeth as $Ca_3(Po_4)_2$

Question Bank:

1. The decrease in melting and boiling points of alkali metals with rise in atomic number is due to
   - (1) the weakening of covalent bonding
   - (2) weak ionic bonding
   - (3) Van der Waal's forces
   - (4) the weakening of the metallic bonding

2. Which of the following set of atomic numbers refers to that of alkali metals
   - (1) 1, 12, 30, 4, 62
   - (2) 37, 19, 3, 55
   - (3) 9, 17, 35, 53
   - (4) 12, 20, 56, 88

3. Regarding the isolation of alkali metals the correct statement is
   - (1) Alkali metals can be isolated by the reduction of their oxides by carbon
   - (2) Alkali metals can be prepared from their aqueous salt solutions by metal displacement method.
   - (3) Alkali metals can be prepared by the electrolysis of their aqueous salt solutions
   - (4) Alkali metals can be prepared by electrolysis of their fused chlorides.

4. Which of the following alkali metal ions in aqueous solution is the best conductor of electricity

   1) Li^+ (2) Na^+

   (3) Cs^+ (4) K^+

5. Alkali metals is each period have
   - (1) smallest size
   - (2) lowest ionisation energy
   - (3) highest ionisation energy
   - (4) highest electronegativity

6. In alkali metals family cesium should be
   - (1) the least Ionisable
   - (2) more electropositive than others
   - (3) lightest of all the metals
   - (4) having low melting and boiling points

7. Which of the alkali metals has the polarising power equal to that of Magnesium

   (1) Rb (2) K

   (3) Na (4) Li

8. The following statement about alkali metals and their compounds is correct
   - (1) cesium is used in photoelectric cells
   - (2) molten sodium chloride on electrolysis give sodium at the anode and chlorine at the cathode
   - (3) alkali metal atom has the smallest size in its period
   - (4) alkali metals do not react with water

9. The nature of Cesium oxide is

   (1) acidic (2) basic

   (3) weakly basic (4) amphoteric

10. The carbonate of the following is insoluble in water

(1) sodium (2) potassium
(3) lithium (4) Cesium

11. The most dangerous method of preparing hydrogen would be by the action of HCl on
(1) Zn (2) K
(3) Fe (4) Al

12. Regarding alkali metals the wrong statement is
(1) Li^+ has higher I.P than helium
(2) Sodium is amphoteric in nature
(3) Lithium is the weak reducing agent
(4) All alkali metal give blue solution in liquid ammonia

13. Regarding alkali metals the correct statement is
(1) they are powerful oxidising agents
(2) they are powerful reducing agents
(3) they have no oxidising or reducing property
(4) they are both oxidising and reducing agents

14. Which of the following is a wrong statement
(1) potassium reacts with water and gives alkaline solution with the liberation of hydrogen gas.
(2) The alkaline metal hydroxides are strong bases because they ionise completely in dilute aqueous solutions
(3) Potassium carbonate cannot be prepared by solvay process
(4) Lithium chloride is insoluble in alcohol, but soluble in water.

15. Alcohol dissolves the following salt
(1) KCl (2) NaCl
(3) LiCl (4) RbCl

16. Which of the following atoms have the smallest size
(1) Mg (2) Na
(3) Be (4) Li

17. The chloride of an element *A* gives a neutral solution in water. In the periodic table, the element *A* belongs to
(1) first group
(2) third group
(3) fifth group
(4) first transition series

18. Strongest bond is in between
(1) CsF (2) NaCl
(3) both 1and 2 (4) none of these

19. Alkali metals are generally extracted by
(1) reduction methods
(2) double decomposition methods
(3) displacement methods
(4) electrolytic methods

20. Which of the following statements is incorrect
(1) sodium is the not most abundant metal in earth's crust
(2) sodium is the most abundant metallic element in sea water

(3) melting and boiling points of alkali metals decrease down the group
(4) ionic character of alkali metal halides decrease down the group

21. Which one of the following has highest electropositive character

(1) Cu (2) Cs
(3) Ba (4) Cr

22. The electronic configuration of metal M is $1s^2$, $2s^2$, $2p^6$, $3s^1$. The formula of its oxide would be

(1) MO (2) M_2O
(3) M_2O_3 (4) MO_2

23. Lattice energy (numerical value) of chloride of alkali metals in order

(1) LiCl>NaCl>KCl>RbCl>CsCl
(2) LiCl<NaCl<KCl<RbCl<CsCl
(3) NaCl<KCl<LiCl<RbCl<CsCl
(4) NaCl<KCl<RbCl<CsCl<LiCl

24. The most stable halide of alkali metal

(1) sodium fluoride
(2) sodium bromide
(3) sodium chloride
(4) sodium iodide

25. Which is paramagnetic

(1) KO_2 (2) K_2O_2
(3) K_2O (4) NO_2

26. Which of the following is not the characteristics of alkali metals

(1) low melting point
(2) low electronegativity
(3) high ionisation energies
(4) their ions are isoelectronic with noble gases.

27. Assertion: Alkali metals do not occur in the free state
Reason: They have high reactivity

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true an **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

28. Assertion: Alkali metals are s-block elements
Reason: The differentiating electron enters into the s-orbital.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true an **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

29. Assertion: Alkali metals are soft
Reason: Weak interatomic forces of attraction are present in alkali metals.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true an **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

30. Assertion: Sodium in liquid NH_3 is a mixed conductor
Reason: Its electrical conductivity is due to presence of ions and solvated electrons
(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true an **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

31. Match the following.

List-I	**List-II**
A) Borax	1) NaCl
B) Saji Mitti	2) $NaNO_3$
C) Mirabilite	3) Na_2CO_3
D) Rock salt	4) Na_2SO_4
	5) $Na_2B_4O_7.10H_2O$

The correct match from List-I and List-II is

	A	**B**	**C**	**D**
(1)	5	3	2	1
(2)	5	3	4	1
(3)	5	2	4	1
(4)	5	3	1	2

32. Match the following.

List-I *Compound*	**List-II** *Flame colouration*
A) Lithium chloride	1) Blue violet
B) Potassium chloride	2) Red violet
C) Caesium chloride	3) Golden yellow
D) Rubidium chloride	4) Lillac
	5) Crimson red

The correct matching

	A	**B**	**C**	**D**
(1)	5	4	2	1
(2)	5	3	2	1
(3)	3	4	2	1
(4)	5	4	1	2

33. Match the following

List-I *Process*	**List-II** *Electrolyte*
A) Nelson cell	1) Fused NaCl
B) Castener's process	2) Brinc solution
C) Down's process	3) Fused NaOH
D) Castener-Kellner cell	4) Fused $NaNO_3$
	5) Cryolite

The correct matching from List-I and List-II is

	A	**B**	**C**	**D**
(1)	2	3	1	4
(2)	3	5	1	2
(3)	2	3	1	2
(4)	1	3	2	5

34. Which of the following reactions, MgO is not formed? (2005 E)
(1) $Mg + CO_2 \rightarrow$
(2) $Mg + dil.HNO_3 \rightarrow$
(3) $Mg + NO \rightarrow$
(4) $Mg + B_2O_3 \rightarrow$

35. Match the following :

List-1	**List-II**
A) Highest abundance	I) Be
B) Least S.O.P.	II) Mg
C) Highest M.P.	III) Sr
D) Electron	IV) Ca
.	V) Ba

	A	**B**	**C**	**D**
(1)	IV	I	V	II
(2)	IV	I	III	V
(3)	II	I	III	IV
(4)	II	III	IV	V

36. Alkaline earth metals are

A) Be and Cs B) Be and Ba

C) Rb and Be D) Sr and Ra

(1) A, B and C are true

(2) B and D are true

(3) C and D are true

(4) A and D are true

37. Identify the correct statements

A) Magnesium is lighter than calcium.

B) The atomic radius of Mg is greater than that of Ca.

C) Mg alloys are used in the construction of aircrafts.

D) Mg is used as a reducing agent

(1) C and D (2) A and D

(3) A, B and C (4) A, B, C & D

38. **A:** Helium and beryllium have similar outer electronic configuration of the type ns^2.

R: Both are chemically inert.

39. Thermal stability of $BaCO_3$. $CaCO_3$ and $MgCO_3$ is

(1) $CaCO_3 > MgCO_3 > BaCO_3$

(2) $MgCO_3 > CaCO_3 > BaCO_3$

(3) $BaCO_3 > MgCO_3 > CaCO_3$

(4) $BaCO_3 > CaCO_3 > MgCO_3$

40. Setting of plaster of paris is involves

(1) oxidation with atmospheric oxygen

(2) combination with atmospheric CO_2

(3) dehydration

(4) hydration to yield another hydrate

41. A radioactive element X decays giving two inert gases is

(1) ${}_{92}U^{238}$ (2) ${}_{88}Ra^{226}$

(3) ${}_{90}{}^{Th234}$ (4) ${}_{93}Np^{227}$

42. $Ca(NO_3)_2$ is known by the name

(1) chile salt petre

(2) norwegian salt petre

(3) Inidan salt petre

(4) nitre

43. With respect to Be and AI which is not correct?

(1) carbides of both on hydrolysis give methane

(2) they form beryllates and aluminates respectively

(3) $Be(OH)_2$ and $AI(OH)_3$ are bases

(4) Be and AI have same atomic size

44. The number of stages involved in the setting of plaster of paris is

(1) 1 (2) 2 (3) 3 (4) 4

45. Calcium slats give a white crystalline precipitate with ammonium oxalate solution. The precipitate is insoluble in

(1) acetic acid (2) dil nitric acid

(3) dil HCI (4) all of these

46. Slaked lime is used in the manufacture of

(1) cement (2) bricks

(3) pigment (4) medicine

47. Lithopone is a mixture of ZnS and
(1) $BaSO_4$ (2) $BaCO_3$
(3) $BaCl_2$ (4) $ZnSO_4$

48. Which of the following is most thermally stable compound?
(1) $BeCO_3$ (2) $CaCO_3$
(3) $CaCO_3$ (4) $BaCO_3$

49. Calcium metal is used to produce high vacuum because it
(1) can remove water
(2) can remove oxygen and nitrogen
(3) is a good reducing agent
(4) it is good drying agent

50. The formula of plaster of paris is
(1) $CaSO_4.2H_2O$
(2) $CaSO_4.H_2O$
(3) $CaSO_4.3H_2O$
(4) $CaSO_4.H_2O$

51. Which of the following has the highest hydration energy?
(1) Be^{2+} (2) Mg^{2+}
(3) Ca^{2+} (4) Sr^{2+}

52. Carnallite is the ore of
(1) Na (2) Ca
(3) Mg (4) Al

53. Portland cementclinker has highest percentage of
(1) dicalcium silicate
(2) tricalcium silicate
(3) dicalcium aluminate
(4) tricalcium aluminate

54. Weight percentage of gypsum in cement is approximately
(1) 2 (2) 5 (3) 9 (4) 15

KEY

1) **4**	2) **2**	3) **4**	4) **3**	5) **2**	6) **4**	7) **4**	8) **1**	9) **2**	10) **3**
11) **2**	12) **2**	13) **2**	14) **4**	15) **3**	16) **4**	17) **1**	18) **2**	19) **4**	20) **4**
21) **2**	22) **2**	23) **1**	24) **1**	25) **1**	26) **3**	27) **1**	28) **1**	29) **1**	30) **1**
31) **2**	32) **1**	33) **2**	34) **2**	35) **1**	36) **2**	37) **1**	38) **3**	39) **4**	40) **4**
41) **2**	42) **2**	43) **4**	44) **2**	45) **1**	46) **4**	47) **1**	48) **4**	49) **2**	50) **4**
51) **1**	52) **3**	53) **2**	54) **1**						

10. IIIA GROUP ELEMENTS

Synopsis:

1. It is called boron family
2. It has 5 elements

B	Boron	5	(He) $2s^2\ 2p^1$
Al	Aluminum	13	(Ne) $3s^2\ 3p^1$
Ga	Gallium	31	(Ar) $3d^{10}\ 4s^2\ 4p^1$
In	Indium	49	(Kr) $4d^{10}\ 5s^2\ 5p^1$
Th	Thallium	81	(Xe) $4f^{14}\ 4d^{10}\ 6s^2\ 6p^1$

3. General electronic configuration is $ns^2\ np^1$
4. Occurrence: Boron occurs as boron, it is rare in occurrence aluminium is the most abundant metal & most abundant III group element _ Ga, In, Tl exist as traces in Zn and Lead Sulphide ores.
5. Metallic nature: Increase from top to bottom, boron is non-metal all are-metals. (white Solids)
6. Atomic size: There is sudden increase from B to Al and then gradual increase is seen due to screening effect.
7. Density: Increases from B to Tl, Al has relatively lower density because of its larger size.
8. Ionization energy: Decreases from B to Al but increases from Al to Tl because in Tl due to less effective shielding of d^{10} electrons the outer electrons are most tightly held. So they are smaller in size than expected.
9. Electropositive character: Increases from Boron to Aluminium and then remains almost constant.
10. Melting Point: M.P of boron is very high, as it exist as giant covalent polymer in liquid and solid forms. M.P decreases from B to Ga and increases for Thalium low M.P of Gallium is due to simple Ga_2 molecules. The melting point trend is B > Al > Tl > In > Ga.
11. Boiling Point: Decreases from B to Th. It changes as B > Al > In > Tl > Ga.
12. Oxidation state: Common oxidation state is +3. Except 'B' other elements show +1 due to inert pair effect stability of +1 state increases from top to bottom B shows –3 in Mg_3B_2.
13. Reactivity: Amorphous B is reactive; Al is stable due to formation of stable oxide layer. Ga, In & Tl are not attacked by water and are stable in air.
14. Boron is diagonally related to silicon due to its small size & same electro negativity.
15. Similarities between Al and B:
    - (i) Both are non metals
    - (ii) Both are bad conductors, have low densities.
    - (iii) They exist as amorphous & crystalline form.
    - (iv) Both form covalent hydrides & halides.
    - (v) Both form oxides on heating with air.
    - (vi) Both $AlCl_3$ & BF_3 are electron deficient compounds.
    - (vii) Both form nitrides which on hydrolysis gives NH_3.

Differences between B and Al:

Boron	**Aluminium**
1) It is a non-metal.	It is a metal.
2) It is a non-conductor.	It is a good conductor.
3) Its co-valency is 4.	It is co-valency is 6.
4) It shows allotropy.	It does not show allotropy.
5) It does not form cation.	It forms cation.
6) It forms acidic oxides	It forms amphoteric oxide.
7) It forms stable borates.	It forms unstable aluminates

16. Oxides:-They form oxides of formula M_2O_3 called sesquioxide.B_2O_3 – acidic, Al_2O_3 and Ga_2O_3 – amphoteric, In_2O_3 and Tl_2O_3 – basic.
17. Hydrides are covalent Al and Ga show an ionic hydrides, $LiAlH_4$ is powerful reducing agent.
18. Halides: General formula of halides is MX_3.Due to small size and all are covalent and Lewis acids. Lewis acid character should follow.$BF_3 > BCl_3 > BBr_3 > BI_3$. But it is $BI_3 > BBr_3 > BCl_3 > BF_3$
19. This order is due to formation of resonating dative pi bond between boron & halogen. This is called back-bonding or P-II - P-II bonding.
20. Tendency of back bonding decreases from BF_3 to BI_3 because overlapping of Boron *p*-orbitals for back bonding does not take place with higher energy level *p*-orbitals.
21. Boron: Borax exist in three forms
    - (i) Prismatic crystals $Na_2B_4O_7\ 10H_2O$
    - (ii) Octahederal crystals $Na_2B_4O_7\ 5H_2O$.
    - (iii) Anhydrous Sodium Borax glass ($Na_2B_4O_7$) it is also called suhaga.
22. Allotropy:-Boron exists as amorphous and crystalline boron.
23. Boranes: Boron can form a series of hydrides called boranes, which are electron deficient compounds. Boranes are of eight types, four with general formula B_nH_{n+4} and four with B_nH_{n+6}. Diborane - B_2H_6, 9, Penta borane - B_5H_9, Hexa borane - B_6H_{10} and 14Decaborane - $B_{10}H_{14}$ belongs to B_nH_{n+4} series. Tetra borane – B_4H_{10}, Ennea borane – B_9H_{15}, 11Penta borane - B_5H_{11}, 16Decaborane – $B_{10}H_{16}$ belongs to B_nH_{n+6} series.
24. B_2H_6 (Diborane): Prep: it is prepared by reduction of boron halides.

    $4BCl_3 + 3LiAlH_4 \rightarrow 2B_2H_6 + AlCl_3 + HCl$

    $BBr_3 + H_2 \rightarrow B_2H_6 + HBr$.
25. Reactions of diborane:
    - With water is gives boric acid
    - With halogen they form halides.
    - With alkali they give metaborate ($NaBO_2$)
    - With NH_3 at low temp it gives diammoniate of diborane at high temp it gives borazole.

    $B_2H_6 + NH_3 \rightarrow B_2H_6\ .2NH_3 \xrightarrow{\Delta} B_3N_3H_6$.
26. Borazole: It is also called as borazine, inorganic benzene or triboro tri ammine. It resembles benzene in structure. Boron exhibits sp^2 hybridization in borazole. The nitrogen atom in borazole will act as Lewis base by forming dative bond with boron.

27. Structure of Diborane (B_2H_6):
   - (i) It is electron deficient.
   - (ii) It has 4 terminal hydrogen & 2bridge hydrogen.
   - (iii) The $2BH_2$ groups are coplanar
   - (iv) It has sp^3 hybridization.
   - (v) It has banana bond or tau bond or 3 centered 2-electron bond
   - (vi) Bond angle of H-B-H bridge is 97°.
   - (vii) Bond angle of H-B terminal hydrogen is 121.5°
   - (viii) Distance between 2 boron atoms is 1.77 A.

28. Uses of Diborane
   - used as rocket fuel in supersonic bombers.
   - in prep of high energy fuels and propellants.
   - in welding torches.
   - as catalyst in polymerization.

1 Borax

It is the most important compound of boron. It is a white crystalline solid of formula $Na_2B_4O_7 \times 10H_2O$. In fact it contains the tetranuclear units $[B_4O_5(OH)_4]^{2-}$ and correct formula; therefore, is $Na_2[B_4O_5(OH)_4].8H_2O$. Borax dissolves in water to give an alkaline solution.

$$Na_2B_4O_7 + 7H_2O \rightarrow 2NaOH + \underset{\text{Orthoboric acid}}{4H_3BO_3}$$

On heating, borax first loses water molecules and swells up. On further heating it turns into a transparent liquid, which solidifies into glass like material known as borax bead.

$$Na_2B_4O_7.10H_2O \xrightarrow{\Delta} Na_2B_4O_7 \xrightarrow{\Delta} \underset{\text{Sodium metaborate}}{2NaBO_2} + \underset{\text{Boric anhydride}}{B_2O_3}$$

The metaborates of many transition metals have characteristic colours and, therefore, borax bead test can be used to identify them in the laboratory. For example, when borax is heated in a Bunsen burner flame with CoO on a loop of platinum wire, a blue coloured $Co(BO_2)_2$ bead is formed.

Orthoboric acid

Orthoboric acid, H_3BO_3 is a white crystalline solid, with soapy touch. It is sparingly soluble in water but highly soluble in hot water. It can be prepared by acidifying an aqueous solution of borax.

$$Na_2B_4O_7 + 2HCl + 5H_2O \rightarrow 2NaCl + 4B(OH)_3$$

It is also formed by the hydrolysis (reaction with water or dilute acid) of most boron compounds (halides, hydrides, etc.). It has a layer structure in which planar BO_3 units are joined by hydrogen bonds.

Boric acid is a weak monobasic acid. It is not a protonic acid but acts as a Lewis acid by accepting electrons from a hydroxyl
ion:

$B(OH)_3 + 2HOH \rightarrow [B(OH)_4]^- + H_3O^+$

On heating, orthoboric acid above 370K forms metaboric acid, HBO_2 which on further heating yields boric oxide, B_2O_3.

$H_3BO_3 \xrightarrow{\Delta} HBO_2 \xrightarrow{\Delta} B_2O_3$

29. Uses of Al:
    - (i) It is used for making electrical cabels.
    - (ii) On account of its lightness and the resistance to corrosion, aluminium is largely used as alloys, in making (a) working utensils (b) trays, picture frames, combs etc.
    - (iii) It is used in steel industry to remove dissolved gases from molten steel and thus prevent the formation of blow holes.
    - (iv) It gives foils of very small thickness which are used as wrappers for chocolates and cigarettes.
    - (v) Fine dust of the metal is used ink the preparation of paints and lacquers (silver paint)
    - (vi) Mixed with iron oxide, aluminium powder forms thermit which is employed for welding purposes (see aluminothermic process)
    - (vii) the flash light bulbs used in door photography contain aluminium powder
    - (viii) A mixture of aluminium powder and ammonium nitrate is called "ammonal". It is used in bombs.
    - (ix) Bright aluminium surfaces are used in telescopes.
    - (x) Aluminium is used in place of zinc and tin for coating surfaces of iron

CHEMICAL PROPERTIES OF 13 GROUP ELEMENTS

1) Action with air (or) O_2: Amosphous B, on heating in air about 700°C form B_2O_3 and BN.

$$4B + 3O_2 \rightarrow 2B_2O_3$$
$$2B + N_2 \rightarrow 2BN$$

2) With Non metals: On heating amorphous B will react with N2, O_2, & F_2, Cl_2, Br_2, directly.

$$4B + 3O_2 \rightarrow 2B_2O_3$$
$$2B + 3S \xrightarrow{1200^0 C} B_2S_3$$

Boron tri sulphide

Hardest compound of Boron is Boron carbide (B_4C)

3) Reducing property: B is a powerful reducing agent for $Co_2 \times SiO_2$

$$3SiO_2 + 4B \rightarrow 2B_2O_3 + 3S$$

4) With conc. Acids B gives boric acid.

$$B + 3HNo_3 \rightarrow H_3Bo_3 + 3No$$
$$2B + 3H_2So_4 \rightarrow 2H_3Bo_3 + 3So_2$$

5) When fused with alkalies, B liberates H_2

$$2B + 6NaoH \rightarrow 2Na_3Bo_3 + 3H_2$$

Uses: B is good neutron absorb B and BC_3 are used in shielding of atomic piles and in

Control rods used for controlling of chain reactions.

- High purity B is used as a semi conductor in the place of Ge & Si.
- Boron is essential minor element for healthy grounds of plants.
- Boric acid is used in the manufacturing of Pyrex glass.
- Borax is used as flux for soldering metal.
- Orthoboric acid used as mild anticeptic.

Question bank:

1. The sudden increase in the atomic radius from B to Al is due to
(1) the presence of vacant *p*-orbitals
(2) the presence of vacant *d*-orbitals
(3) the presence of lower I.P values
(4) the greater screening effect of electrons in penultimate shell

2. Aluminia is ----- in nature
(1) acidic (2) basic
(3) amphoteric (4) neutral

3. Which of the following does not belong to IIIA group
(1) B (2) Al (3) Ge (4) In

4. The product formed in the reaction $BCl_3 + H_2O \rightarrow$ product is
(1) $H_3PO_3 + HCl$ (2) $B_2O_3 + HOCl$
(3) $B_2H_6 + HCl$ (4) no reaction

5. The boron, carbon, nitrogen and oxygen have their increasing ionisation energy in the order
(1) B, C, O, N (2) N, O, B, C
(3) B, N, C, O (4) O, B, C, N

6. Which of the following is not a Lewis acid
(1) SiF_4 (2) $FeCl_3$
(3) BF_3 (4) C_2H_4

7. $AlCl_3$ exists as dimer because
(1) Al has greater I.P
(2) Al has larger radius
(3) High charge nucleus
(4) Incomplete *p*-orbitals

8. The non metal is
(1) indium (2) boron
(3) aluminium (4) gallium

9. Generally, the atomic and ionic radii should increase with increasing atomic number down the group. But there is no much increase at Ga and Tl in group 13 and it is due to
(1) the high nuclear charge of Ga and Tl
(2) poor shielding d and f electrons
(3) high ionisation energies of these elements
(4) none of these

10. III A elements exhibit
(1) only +1 oxidation state
(2) only +3 oxidation state
(3) both +1 and +3 oxidation state
(4) +1, +2 and +3 oxidation states

11. Which of the following statements is not true for both Al and B
(1) they burn in oxygen to give oxides at higher temperature
(2) their halides are Lewis acids
(3) they combine with nitrogen to form nitrides

(4) they react with HCl to give chloride

12. Monovalancey of heavier members of group can be explained on the basis of
(1) their low ionisation energies
(2) their low electronegativity values
(3) their larger sizes
(4) inert pair effect

13. When aluminium is heated in a current of nitrogen gas, a compound *X* is formed. Then the oxidation state of *N* in *X* is
(1) +3 (2) –3 (3) +5 (4) +4

14. The primary and secondary valencies of aluminium in meta aluminate or aluminate ions are
(1) 1 and 6 (2) 2 and 6
(3) 3 and 3 (4) 3 and 6

15. Which is the aluminate ion
(1) $[Al(OH)_4(H_2O)^-]$
(2) $[Al(OH)_3(H_2O)_3]$
(3) $[Al(OH)_6]^{3-}$
(4) $[Al(OH)_5(H_2O)]^{3-}$

16. Which is the correct representation of meta aluminate ion
(1) $[Al(OH)_3]^-$
(2) $[Al(OH)_3(H_2O_3)]^{3-}$
(3) $[Al(OH_4)(H_2O_2)]^-$
(4) $[Al(OH_2)(H_2O_4)]^+$

17. Aluminium does not react with
(1) pure cold water
(2) conc HNO_3
(3) dry air
(4) all of these

18. Aluminium containers are used to transport
(1) soda lye (2) conc H_2SO_4
(3) conc HNO_3 (4) conc HCl

19. Observe the following reaction $Al + \text{conc } H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O + X$. Regarding *X* the wrong statement is
(1) the central atom in *X* is sp^2 hybridised
(2) the shape of *X* is trigonal planar
(3) it has a lone pair of electrons
(4) it acts as a reducing agent and oxidising agent

20. $Al + \text{dil } HNO_3 \rightarrow Al(NO_3)_3 + X + H_2O$. The types of bonds present in *X* are
(1) Ionic and covalent
(2) Ionic and dative
(3) covalent and dative
(4) Ionic, covalent and dative

21. Which of the following does not react with water
(1) boron (2) aluminium
(3) sodium (4) thallium

22. The following property changes in a regular manner from B to Tl
(1) (1)melting point
(2) ionisation potential
(3) boiling point
(4) density

23. Backbonding of electrons pairs of Pz-orbital of central atom is more promient in
(1) BF_3 (2) BCl_3
(3) BBr_3 (4) BI_3

24. Anhydrous $AlCl_3$ is a catalyst in
(1) Wurtz reaction
(2) Friedel-Craft's reaction
(3) Willamson synthesis
(4) Haeber's process

25. $2M + 6HCl \rightarrow 2MCl_3 + 3H_2$. From this reaction, what would be M from the following which does not displace hydrogen from hydrochloric acid

(1) B (2) Al (3) Ga (4) Tl

26. Assertion: Boron exhibit allotropy

Reason: Boron exists in crystalline and amorphous forms

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

27. Assertion: Lithium does not form alum.

Reason: The size of Lithium is smaller and can not accommodate 6 water molecules

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

28. Assertion: The nature of aqueous potash alum solution is acidic

Reason: It is due to the hydrolysis of $Al_2(SO_4)_3$.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

29. The type of bonds present in the dimer of aluminium chloride are

(1) 4 covalent bonds, 4 dative bonds
(2) 5 covalent bonds, 3 dative bonds
(3) 2 covalent bonds, 6 dative bonds
(4) 6 covalent bonds, 2 dative bonds

30. Assertion: Anhydrous $AlCl_3$ is a covalent compound.

Reason: The ionisation energy of Al is more

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

31. Assertion: +1 oxidation state is more prominent in Thallium than in Gallium

Reason: The inert pair effect is more in Thallium than in Gallium

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

32. Some ores of aluminium were mentioned.

(a) corundum (b) cryolite
(c) gibbsite (d) bauxite
(e) diaspore (f) feldspar
(g) awnite

(1) a, b, c & d (2) b, d & g
(3) c & d (4) e, f & g

33. Pick out the correct statements.
(a) The percentage purity of Al obtained from Hall-Heroult's process is 99.98%
(b) In Hoope's process the carbon rod acts as cathode
(c) All forms amalgam and alloys with other elements
(d) It is difficult to solder Al

The correct answer is
(1) b, c & d (2) a, b & d
(3) b & d (4) all are right

34. The substance which is used in chromatography technique is
(1) $AlCl_3$ (2) Al_2O_3
(3) $Al_2(SO_4)_3$ (4) AlN

35. Precius ruby stones are
(1) aluminium silicate
(2) alumina
(3) sodium aluminium silicate
(4) sodium silicate

36. Boric Acid is a
(1) A weak mono basic acid
(2) Protic acid
(3) Di Protonic acid
(4) Try Protonic acid

37. White fumes appear around the bottle of anhydrous almuminum chloride is due to
(1) Moist HCl gas
(2) Moist NH_3 gas
(3) Moist NH_4Cl gas
(4) Moist $AlCl_3$ gas

38. Match the following.

Reactants	**Products**
A) $B_2H_6 + Br_2$	1) $B_2H_5Cl + H_2$
B) $B_2H_6 + Cl_2$	2) $B_2H_5Cl + HCl$
C) $B_2H_6 + HCl$	3) $BCl_3 + HCl$
D) $B_2H_6 + CO$	4) BH_3CO
	5) $B_2H_5Br + HBr$

The correct match is

	A	**B**	**C**	**D**
(1)	1	3	2	4
(2)	5	3	2	1
(3)	5	3	1	4
(4)	2	4	3	1

39. Match the following.

List – I	**List – II**
A) Boron	1) Low I.P.
B) Aluminium	2) Low M.P.
C) Gallium	3) Maximum covalency 4
D) Thallium	4) Higher density
	5) Maximum covalency 6

The correct match is

	A	**B**	**C**	**D**
(1)	1	2	3	4
(2)	3	5	2	4
(3)	5	1	4	2
(4)	1	3	5	4

40. Match the following.

Reaction of Al with	**One of the products formed**
A) air	1) Al_2O_3
B) conc. H_2SO_4	2) $NaAlO_2$
C) fused NaOH	3) SO_2
D) aqueous NaOH	4) Na_3AlO_3
	5) $Al(NO_3)_3$

The correct match is

	A	**B**	**C**	**D**
(1)	3	4	2	5
(2)	1	2	3	4
(3)	5	4	2	3
(4)	1	3	4	2

1) **4** 2) **3** 3) **3** 4) **1** 5) **1** 6) **4** 7) **4** 8) **2** 9) **2** 10) **3**
11) **4** 12) **4** 13) **2** 14) **4** 15) **3** 16) **3** 17) **4** 18) **3** 19) **2** 20) **4**
21) **4** 22) **4** 23) **1** 24) **2** 25) **1** 26) **1** 27) **1** 28) **1** 29) **4** 30) **1**
31) **1** 32) **2** 33) **1** 34) **2** 35) **3** 36) **1** 37) **1** 38) **3** 39) **2** 40) **4**

11. IVA GROUP ELEMENTS

Synopsis:

1. The IV group elements are Carbon (C), Silicon (Si), Germanium (Ge), Tin (Sn) and Lead (Pb).
2. The general valence shell configuration of IVA elements is $ns^2\ np^2$.
3. The electron configuration is

Elements	Symbol	Electron Configuration
Carbon	C	(He) $2s^2 2p^2$
Silicon	Si	(Ne) $3s^2 3p^2$
Germanium	Ge	(Ar) $3d^{10} 4s^2 4p^2$
Tin	Sn	(Kr) $4d^{10} 5s^2 5p^2$
Lead	Pb	(Xe) $4f^{14} 5d^{10} 6s^2 6p^2$

4. The most abundant element of this group is Si and the least abundant element is Ge.
5. The order of melting points of IVA elements is: C >Si > Ge > Pb > Sn
6. The order of boiling points of IVA elements is: C > Ge > Sn > Si > Pb
7. The atomic size increases from C to Pb.
8. The difference in the atomic size of "Si" and "Ge" is less. It is due to ineffective shielding of nuclear charge by the completely filled 3*d*-shell in "Ge".
9. The difference in the atomic size of "Sn" and "Pb" is also very less. It is due to the presence of completely filled 4*f*-orbitals in "Pb".
10. The density decreases from C to Si, and then increases from Si and Pb.
    Pb > Sn > Ge > C > Si
11. Ionisation potentials decrease from C to Pb.
12. Electronegativity of Carbon is 2.1. Electronegativity of Si, Ge, Sn and Pb are equal to1.8.
13. Carbon shows a large no of oxidation states, the other members of the family exhibits +2 and +4 oxidation states.
14. Sn^{+4} is more stable than Sn^{+2} but Pb^{+2} is more stable than Pb^{+4}.
15. The linkage of atoms of the same element to form long chain is called catenation.
16. Catenation ability is highest for carbon. This is due to its high bond strength.
17. Order of bond energies: C-C > Si-Si > Ge- Ge > Sn –Sn > Pb- Pb.
    Order of catenation ability: C > Si > Ge > Sn > Pb.
18. "C" and "Si" exhibit tetravalency. "Si" exhibits maximum covalency of 6.
19. "C" and "Si" non-metals and can form hydrides, halides and oxides.
20. CO_2 and SiO_2 are acidic oxides.
21. Carbon is the essential element of plant and animal kingdom where as Si is essential element of mineral kingdom.
22. Diamond the crystalline allotrope of carbon, the hardest naturally occurring substance.
23. Among the crystalline allotropes of carbon, the good conductor of electricity is graphite. Among amorphous allotropes of carbon, good electrical conductors are gas carbon and petroleum coke. Si is a semiconductor.
24. Fullerenes consist of C_{60}. It is also called Buckminsterfullerene.

25. It is Soccer ball like structure consist of 12 pentagonal rings and 20 hexagonal rings. The carbon atom is sp^2 hybridised.
26. Due to the high catenation ability, Carbon forms millions of compounds.
27. Carbon can form multiple bonds. Generally silicon forms single bonds.
28. Alkanes alkenes and alkynes are the hydrides of carbon. They do not undergo hydrolysis. They are stable.
29. Hydrides of silicon are called silanes.
30. In Silanes, the electronegativity of hydrogen is more than that of “Si”. Hence Si-H and Si- Si bonds are weakened. Silanes are unstable and can undergo hydrolysis.
31. Silances are good reducing agents due to weaker Si-Si and Si-H bonds
32. C and Si form tetrachlorides, which are covalent. They have tetrahedral structures.
33. CCl_4 does not undergo hydrolysis due to the absence of *d*-orbitals in carbon.
34. $SiCl_4$ undergoes hydrolysis due to the presence of vacant 3*d*-orbitals in “Si” atom.
35. Oxides of carbon are CO and CO_2. Silicon forms only one oxide--- SiO_2
36. CO_2 molecule contains two double bonds. It exists as discrete molecules. Hence it is a gas.
37. SiO_2 has giant molecular structure and exists as solid.
38. In SiO_2, each silicon atom is tetrahedral surrounded by 4 oxygen atoms. Its geometry is similar to that of diamond.
39. The boiling point of CO_2 is very low since very weak attractive forces exist between the molecules of CO_2.
40. SiO_2 has giant three dimensional network structure. Hence its m.pt. is high.
41. Existence of an element in more than one physical form is known as allotropy. Allotropes of an element differ to a greater extent in their physical properties but exhibit close resemblance in the chemical properties.
42. Crystalline allotropes of carbon are--- Diamond, Graphite and Fullerene
43.

Diamond	**Graphite**
a) Each carbon is sp^3 hybridised.	1) Each carbon is sp^2 hybridised.
b) Each carbon is bonded to 4 other carbons tetrahedrally	2) Each carbon is bonded to 3 other atoms to form hexagonal rings. It has sheet like structure.
c) It is a 3 dimensional polymer.	3) It is a 2 dimensional polymer.
d) C-C bond length is 1.54A° and bond angle is 109° 28^1	4) C-C bond length in hexagonal rings is 1.42A° and bond angle is 120°.
e) Carbon atoms are firmly held with strong covalent bonds.	5) The distance between two adjacent layers is 3.35A°. These layers are held by weak vander wall’s forces.
f) Diamond is very hard	6) Graphite is soft.
g) Density = 3.5 g/cc	7) Density = 2.2 g/cc
h) Diamond is an insulator due to the absence of free electrons.	8) Graphite is a conductor due to the presence of free electrons.
i) It is transparent to light and -rays. It has high refractive index (2.45).	9) It, has layer lactice. The layers X are slippery. Hence it is greasy.

44. The weight of diamond is expressed in carat.1 carat= 0.2 g = 200 mg. The allotropes of carbon are insoluble in any solvent.

45. Melting point of Diamond is 4200K. It is stable upto 1500°c in vaccum and thereafter gets converted into Graphite.

$$\text{Diamond} \xrightarrow{\substack{1800\text{-}2000C \\ \text{Vaccum}}} \text{Graphite}$$

46. Uses of Diamond: Diamonds are used
    a) as precious stones in jewellery.
    b) for drawing thin wires.
    c) for cutting glass, drilling rocks and as abrasives.
47. Graphite is manufactured by Acheson's process
    Acheson's process:-
    $3C + SiO_2 \rightarrow SiC + 2CO$
    $SiC \rightarrow \underset{\text{(graphite)}}{C} + Si$
48. Graphite is thermodynamically more stable than diamond.
49. Graphite when heated in the absence of air or Oxygen, it gives CO_2.
50. Layers of graphite are held together by vanderwall's forces.
51. Uses of Graphite:
    (i) as a lubricant.
    (ii) for making electrodes in electrical furnaces
    (iii) in electroplating and in electrotyping
    (iv) in the manufacture of Lead pencils.
    (v) Graphite is used in the manufacture of refractory crucibles.
52. Amorphous allotropes of carbon are --- coal, coke, wood charcoal, animal charcoal, lamp black, carbon black, gas carbon, petroleum coke and sugar charcoal.
53. Silicon Dioxide is called silica. Silica has two allotropic forms namely crystalline and amorphous forms.
54. Crystalline forms of silica are: 1. Quartz 2. Tridymite and 3. Crystobalite
55.
    (i) The purest form of silica is quartz or rock crystal
    (ii) Pure Quartz is colourless crystal. Crushed form of quartz is called sand.
    (iii) Sand particles bonded to iron oxide is known as sandstone.
    (iv) The remains of minute sea organism called Keisulguhr consist of siliceous rock.
56. Effect of temperature on quartz:

α- quartz $\qquad$ α- tridymite $\qquad$ α- cristobalite
$\uparrow\downarrow$845K $\qquad$ $\uparrow\downarrow$400-435K $\qquad$ $\uparrow\downarrow$475-550K
β-quartz1140β-tridymite1740 β-cristobalite
$\uparrow\downarrow$1985K
liquid silica

57. The gem stones agate, onyx, jasper are amorphous forms of silica
58. Preparation:
    a) Silicon when burnt in air or oxygen gives SiO_2. $Si+O_2 \rightarrow SiO_2$
    b) On hydrolysis $SiCl_4$ or SiF_4 gives SiO_2
    $SiCl_4+4H_2O \rightarrow 4HCl + H_4SiO_4$
    $H_4SiO_4 \rightarrow SiO_2 + H_2O$

59. Properties of SiO_2: SiO_2 is insoluble in water.

60. Silica melts at 1600° C to form Quartz glass. Red hot Quartz glass when plunged into water, does not break into pieces. Hence it is used in making optical instruments.

$SiO_2 + HCl.HBr.HI \rightarrow$ no reaction

$SiO_2 + HF \rightarrow SiF_4 + H_2O \rightarrow H_2SiO_4 + H_2SiF_6$

$SiO_2 + NaOH \rightarrow Na_2SiO_3 + H_2O$

$SiO_2 + Na_2O \rightarrow Na_2SiO_3$

$SiO_2 + CaO \rightarrow CaSiO_3$

$SiO_2 + Na_2CO_3 \rightarrow Na_2SiO_3 + CO_2$

$SiO_2 + CaCO_3 \rightarrow CaSiO_3 + CO_2$

$SiO_2 + C \rightarrow SiC+CO$

61. Uses of Silica:
    (i) Silica is used as a building material.
    (ii) Quartz glass is produced by melting silica.
    (iii) Bricks made from a mixture of powdered sand and clay are used for preparing refractory bricks to constant Furnaces.
    (iv) Pure Quartz is used for the manufacture of lenses, optical instruments.
    (v) Coloured Quartz is used as gems.
    (vi) Silica is used as acidic flux in metallurgy.
    (vii) Silicon carbide is carborundum and is used as refractory materials.

62. Structure of Silica: Silica is a giant molecule with 3 dimensional structure. Hence it exists as solid.

63. In SiO_2, each “Si” atom is tetrahedral surrounded by 4 oxygen atoms.

64. Si-O-Si bonds are weaker than C-C bonds in diamond. Hence the m.p. of silica is much less than that of Diamond.

65. CO_2 is acidic oxide. SiO_2 is weekly acidic

66. $SiO_2 + NaHCO_3$ is used as fire extinguisher.

Uses of oxides of carbon:

- CO is a Good reducing agent.
- Carbon mono oxide is used as ligand in metal carbonyl.
- CO_2 is extensively used to carbonate soft drinks.
- Dry Ice (solid CO_2) is used as refrigerant for ice cream and frozen food
- CO_2 is used in Manufature of urea & increasing inert atmosphere.
- CO_2 is used as fire extinguisher.

Silicones

- Silicons are organs – silicon compounds.
- Silicones have silicon strongly cross linked to oxygen and carbon.
- Hydrolysis of chlorosilanes to give polymers.
- Silicons are used to prepare water proof clothes and papers.
- Silicones are used to prepare the grease lubricants in aeroplanes.

Silicates

- Many building materials are called silicates.
- Ceramics and glass are also silicates.

- The Si-O bonds in silicates are very strong.
- Silicates do not dissolve in any of the solvents.
- Silicates consist of SIO_4^{4-} unit which are join by 1.2.3. or 4 silicate unit to form. Chain, sheet or 3 dimentional structure.
- Silicates consist of SIO_4^{4-} unit in tetrahedral shape.
- Negatived charge of silicate structure is neutralized by possitvely charge metal ion.
- Zeolite, feldspar, mica and asbestos are natural silicates.
- Glass and cement are man made silicates.

Zeolites: Zeolites act as ion exchangers and as molecular sieves.

- The structure of zeolites permit the formation of cavities of different sizes.
- In Zeolite some Almunium atoms are replaced by Cation such as Na^+, K^+ or Ca^{2+} which balance negative charge.
- ZSM -5 (A type of zeolite) used to convert alcohols directly into gasoline.

Question Bank:

1. Element of group 14 have outer electronic configuration of
(1) ns^2np^2 (2) ns^2
(3) ns^2np^6 (4) ns^2np^4

2. The electronic configuration of two neutral atoms (*A* and *B*) are given below $A = 1s^22s^22p^2$ and B = [Ne] $3s^23p^2$. The nature of bond between *A* and *B* is mostly
(1) covalent (2) electrovalent
(3) coordinate (4) metallic

3. Which of the following does not exist
(1) CCl_4 (2) SiF_4
(3) $PbBr_4$ (4) $SnCl_4$

4. $SiCl_4$ is easily hydrolysed but CCl_4 is not. This is because
(1) bonding in $SiCl_4$ is ionic
(2) silicon is non-metallic
(3) silicon can extend is coordination number beyond four but carbon cannot
(4) silicon can form hydrogen bonds but carbon cannot

5. The correct increasing order of extent of hydrolysis in the following is
(1) $CCl_4 < MgCl_2 < AlCl_3 < SiCl_4 < PCl_5$
(2) $CCl_4 < AlCl_3 < MgCl_2 < PCl_5 < SiCl_4$
(3) $AlCl_3 < MgCl_2 < CCl_4 < PCl_5 < SiCl_4$
(4) $SiCl_4 < MgCl_2 < AlCl_3 < PCl_5 < CCl_4$

6. The C-C bond energy is 348k J/mole. Then the bond energy of Si-Si is ------
(1) 180 (2) 348
(3) 540 (4) 760

7. The nature of bonds in compounds of C and Si are
(1) electrovalent
(2) covalent
(3) metallic
(4) covalent and electrovalent.

8. Sindhur used by women is an oxide of lead with the formula
(1) PbO (2) PbO_2
(3) Pb_3O_4 (4) Pb_2O_3

9. Which of the following is not a transition metal
(1) chromium (2) titanium
(3) lead (4) tungsten

10. C and Si belong to IV group. The maximum coordination number of carbon is commonly occurring compounds is 4. Whereas that of silicon is 6. This is due to
(1) large size of silicon
(2) availability of low lying d-orbitals in silicon
(3) more electropositive nature of silicon
(4) silicon being vulnerable to attack by nucleophilic reagents

11. The correct order of the size is
(1) Be > C > F > Ne
(2) Be > C > Ne
(3) Be > C > F < Ne
(4) F < Na < Ba < C

12. Which of the following metals shows allotropy
(1) Sn (2) Ca
(3) K (4) Pb

13. Carbon differs from other elements of the group. Which is the false statement
(1) due to limitation of co-ordination number 4
(2) due to d-orbitals in penultimate shell
(3) due to its unique ability to form long chains
(4) due to its marked tendency to form long chains (catenation)

14. Carbon atoms in diamond are bonded with each other in a configuration
(1) Linear (2) Tetrahedral
(3) Planar (4) Octahedral

15. The ability of a substance to assume two or more crystalline structure is called
(1) isomerism (2) polymerism
(3) isomorphism (4) amorphism

16. Wood charcoal is used to decolourise sugar because it
(1) absorbs coloured material
(2) absorbs decolourised material
(3) reduces coloured material
(4) none of these

17. In graphite C undergoes the hybridisation
(1) sp^3 (2) sp
(3) sp^2 (4) dsp^2

18. The purest form of carbon is
(1) diamond (2) graphite
(3) suar charcoal (4) fullerences

19. The allotropic form of carbon soluble in toluene is
(1) diamond (2) graphite
(3) fullerences (4) none

20. Which one of the following has maximum bond dissociation energy
(1) O-O (2) C-C
(3) N-N (4) F-F

21. The element which occurs both in very hard and very soft forms is
(1) Si (2) C
(3) Sn (4) Pb

22. The common semiconductor material is
(1) Fe (2) Se

(3) Ge (4) C

23. Which property is not exhibit by carbon in its compounds
(1) forming bonds to other carbon atoms
(2) forming multiple bonds
(3) exhibiting allotropic forms
(4) forming compounds with coordination number beyond four

24. Diamond and graphite both are made of carbon atoms. Diamond is extremely hard whereas graphite is soft. This is because
(1) the chemical bonds between any two carbon atoms in diamond are stronger
(2) diamond is ionic, whereas graphite is covalent
(3) each carbon atom in diamond is chemically bonded to a greater number of neighboring carbon atoms
(4) certain atoms in diamond are smaller in size

25. Graphite is
(1) harder than diamond
(2) is a bad conductor of electricity
(3) used as a lubricant
(4) an amorphous allotropic form of carbon

26. Which of the following is more stable
(1) Sn^{4+} (2) Pb^{4+}
(3) Sn^{2+} (4) Ge^{2+}

27. Purest form of carbon is
(1) animal charcoal
(2) sugar charcoal
(3) petroleum coke
(4) lamb black

28. The most reactive form of carbon is
(1) diamond (2) graphite
(3) coal (4) charcoal

29. Animal charcoals used to decolourisation of sugar because
(1) it oxidises colured material
(2) it reduces coloured material
(3) it converts coloured material into colourless material
(4) it absorbs coloured material

30. Wood charcoal is used in gas masks because it
(1) is poisonous
(2) is porous
(3) absorbs gases
(4) liquefies gases

31. Graphite conducts electricity because of
(1) weak vander Waals forces between layers
(2) covalent bonding between carbon atoms of layers
(3) delocalised electrons in each layer
(4) sp^2 hybridisation of carbon atom in each layer

32. Which property is common in diamond and graphite
(1) electrical conductivity
(2) relative atomic weight
(3) crystal structure
(4) density

33. In graphite, electrons are
(1) localised on each carbon atom
(2) spreadout between the sheets
(3) localised on every third carbon atom
(4) present in antibonding orbital

34. Assertion: Pb^{4+} can be reduced easily to Pb^{2+}

Reason: Pb^{2+} is paramagnetic.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

35. Assertion: The electronegativity of IVA group elements is same, except for carbon.

Reason: Except carbon and silicon, other elements have completely filled d and f orbitals.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

36. Assertion: $SiCl_4$ undergo hydrolysis easily than CCl_4.

Reason: Silicon has vacant d-orbitals and can accommodate electrons from water molecules.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

37. Assertion: Diamond is a three dimensional polymer of carbon atoms.

Reason: The hybridisation of carbon in diamond is sp^3.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

38. Assertion: Diamonds are used for drilling rocks and cutting glass.

Reason: Diamonds are the hardest substances.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

39. Assertion: Graphite is used in the manufacture of refractory crucibles.

Reason: Graphite is thermally less stable than Diamond.

(1) Both **A** and **R** are true and **R** is the correct explanation of **A**
(2) Both **A** and **R** are true and **R** is not the correct explanation of **A**
(3) **A** is true but **R** is false
(4) **A** is false but **R** is true

40. Which of the following is not the characteristic property of carbon

(1) it exhibits catenation
(2) it forms compounds with multiple bonds
(3) its melting point and boiling point are exceptionally high
(4) it shows semi-metallic character

41. An example of major air pollutant is
(1) O_2 (2) CO_2
(3) CO (4) He

42. NO_2 is released by heating
(1) $Pb(NO_3)_2$ (2) KNO_3
(3) $NaNO_2$ (4) $NaNO_3$

43. Identify the correct statement with respect to carbon monoxide
(1) it combines with water to form carbonic acid
(2) it reacts with haemoglobin in red blood cells
(3) it is a powerful oxidising agent
(4) it is used to prepare aerated drinks

44. Incomplete combustion of petrol or diesel in automobile engines can be best detected by testing the fuel gases for the presence of
(1) carbon monoxide and water vapour
(2) carbon monoxide
(3) nitrogen dioxide
(4) sulphur dioxide

45. Which is a polar bond?
(1) C-H (2) Si-H
(3) C-C (4) Si-Si

46. Which of the following combines with haemoglobin of the blood to form carboxy haemoglobin
(1) CO (2) CO_2
(3) O_2 (4) N_2

47. Percentage of lead in lead pencil is
(1) Zero (2) 20
(3) 80 (4) 70

48. Match the following

Silica reacts with	***Products formed***
A) NaOH	1) $Na_2SiO_3 + CO_2$
B) CaO	2) $CaSiO_3$
C) Na_2CO_3	3) SiC + CO
D) C	4) no product
	5) Na_2SiO_3+H_2O

The correct match is

	A	**B**	**C**	**D**
(1)	4	3	2	1
(2)	5	2	1	3
(3)	3	4	2	5
(4)	1	5	3	4

49. Match the following:

List-I ***Silica form***	**List-II** ***Temperature***
A) α-Quartz	1) 385K
B) α-Tridymite	2) 475-550K
C) α-Crystobalite	3) 400-435K
	4)845K
	5)1985K

The correct match is

	A	**B**	**C**
(1)	5	4	1
(2)	4	3	2
(3)	1	2	3
(4)	1	5	4

50. Match the following:

List-I	**List-II**
A) Water gas	1) Blue gas
B) Producer gas	2) lower calorific value
C) Semi water gas	3) $CO + H_2 + N_2$
D) Coal gas	4) $CO_2 + N_2 + H_2$
	5) highest calorific value

51. The bond length of C-C of fullerene is
(1) 143.5 and 138.3 pm
(2) 138.3 pm and 143.5

(3) 143.5 and 143.5 pm
(4) None of theses

52. Two manmade silicates are
(1) Zeolite and feldspar
(2) Mica and Asbastos
(3) Glass and cement
(4) None of these

The correct match is

	A	**B**	**C**	**D**
(1)	1	2	3	5
(2)	4	2	3	1
(3)	5	2	4	3
(4)	3	5	2	1

KEY

1) **1**	2) **1**	3) **3**	4) **3**	5) **1**	6) **1**	7) **2**	8) **3**	9) **3**	10) **2**
11) **3**	12) **4**	13) **2**	14) **2**	15) **2**	16) **1**	17) **3**	18) **4**	19) **3**	20) **2**
21) **2**	22) **3**	23) **4**	24) **3**	25) **3**	26) **3**	27) **2**	28) **4**	29) **4**	30) **3**
31) **3**	32) **2**	33) **2**	34) **3**	35) **1**	36) **1**	37) **1**	38) **1**	39) **3**	40) **4**
41) **3**	42) **1**	43) **2**	44) **2**	45) **2**	46) **1**	47) **1**	48) **2**	49) **2**	50) **1**
51) **1**	52) **3**								

1) Organo silicon polymers containing the type of linkages is
(1) Si – Si – Si (2) Si – O – Si (3) Si – C- Si (4) None

2) Hydrolysis of chlorosilanes give
(1) Monomers (2) Dimers (3) Polymers (4) Trimers

3) Silicones are
(1) Toxic (2) Si – O – Si (3) Si – C- Si (4) None

4) Which catalyst is employed when alkyl halides when passed over silicon at 300°C to give
(1) Fe (2) Cu (3) Mg (4) Pd

5) Ceramics and glass are also called as
(1) Silicones (2) Zeolites (3) Silicates (4) Insulators

6) Example of pyro silicate is
(1) Throtvetite (2) Willemite (3) Spodumene (4) Kaoline

7) The empirical formula of sheet silicates is
(1) $(Si_2O_5)_n^{2n-}$ (2) $(SiO_3)_n^{2n-}$ (3) $(Si_2O_5)_n^{2n-}$ (4) $(Si_2O_5)_n^{2n-}$

8) The following type of silicate are used for electrical insulation.
(1) Sheet silicate (2) cyclic silicates (3) chain silicates (4) Frame work silicates

9) Zeolites act as
(1) Atomic sieves (2) Molecular sieves (3) Ionic sieves (4) radical sieves

10) In softening of water when zeolites are excersied the following metal ion is replaced by Na^+
(1) Ca^{+2} (2) Ba^{+2} (3) Be^{+2} (4) Zn^{+2}

KEY

1) **2** 2) **3** 3) **2** 4) **2** 5) **3** 6) **1** 7) **1** 8) **1** 9) **1** 10) **1**

13. ENVIORNMENTAL CHEMISTRY

Synopsis:

1. Science dealing with chemical changes in the environment is called environmental chemistry.
2. Environment indicates the surroundings like air water soil, sunlight etc.
3. environment has abiotic & biotic components
4. Abiotic component has atmosphere, lithosphere & hydrosphere
   - Atmosphere – covers the air (gases) surrounding the earth.
   - Lithosphere includes outer month of solid earth with minerals, soils etc. It includes the mountains hills etc.
   - Hydrosphere constitutes all types of water resources like seas, rivers, oceans, lakes, ponds, glaciers etc.
5. Atmoshpere can be divided into 4 major regions.
   i) Troposphere:- it extends from 0-11km above the earth. It is homogenous has air. It contains N_2,O_2, water vapour as main components Ar,CO_2, are minor component, He, Ne, Kr, Xe, NH_3, NO, O_3, SO_2 as trace components. Temperature increases with height.
   ii) Stratosphere: - It extends form 11-50 km. it contains ozone & absorbs U.V radiations.
   iii) Mesosphere: - It extends from 50-85 km temp falls with height sound waves are not propagated (O_2^+; NO^+ are present)
   iv) Thermosphere:- It starts from 85km & extends to 500km. It has maximum attainable temp of 1476k.O_2 undergoes ionization in this zero & forms O_2^+, O^+,NO^+.
6. The energy component includes solar energy, geochemical thermo electrical, hydro-electrical & nuclear energy.
7. Pollution:- the phenomenon by which a substance already present in atmosphere increases then a natural abundance due to human activity & has detrimental effect on the environment & mankind is called pollution.
8. Pollutant :- Substance whose concentration increase in the environment causes adverse effect on mankind or life on earth Eg:- SO_2, CO, NO etc.
9. Contaminant: - This substance is not present in the atmosphere originally but is added due to human activity. A contaminant is called a pollutant when it has detrimental effect. Eg:- Methyl isocyanate (CH_3NCO)
10. Receptor: - Medium that is affected by pollutant. Eg : Man, trees, buildings.
11. Sink:- Medium reacting with pollutant Eg:- Micro organisms, sea
12. Speciation: - Chemical form of pollutant.
13. DO (Dissolved oxygen):- Optimum value for D.O is 4-6mg/lt. Lower D.O value indicates water pollution.
14. COD (Chemical oxygen demand):- The quantity of oxygen required for oxidation of organic compounds is called COD. The COD can be estimated by acidified potassium dichromate.
15. BOD:- The quantity of oxygen used by the waste materials in water in 5 days at 20^0c is called biological oxygen demand. BOD of pure water is 1-3 ppm. The BOD of domestic sewage is 165ppm.

16. TLV:- The permissible level of the toxic pollutant in atmosphere to which a healthy industrial workers is exposed during an eight hour day without any adverse effect is known as threshold limit value (TLV).
17. Common air pollutants are CO, N_2 (Oxides of N_2) oxides of sulphur, hydrocarbons, ozone, CFC, particulates, metals, organic pollutants.
18. Air pollution is mainly due to deforestation, over population, industrialization, transportation (usage of automobiles).

POLLUTANT	SOURCE	HARMFUL EFFECTS
CO	Incomplete combustion of fuels. [acceptable level of CO is 9 ppm]	Forms car boxy hemoglobin -restricts transport of O_2. - High conc. causes death. -Loss of vision, dizziness. -headache, fatigue. -unconsciousness. -leaf drop, leaf curl, reduction in leaf size.
CO_2	Burning of fossil fuels falling of trees.	-green house effect -global warming -cyclones
Hydrocarbons	Produced by natural gas, degradation of organic matter & biomass.	Cause photochemical smog. (Benzopyrene is carcinogenic) PAN, PBN damage plants. Irritation to eyes.
Chloro fluoro carbons	Released from jet planes, air conditioning & refrigerator gases.	Caused depletion of ozone damage of ozone layer causes skin diseases (skin cancer) damages eyes.(cataract of eye)
Oxides of N_2	Combustion of fossil fuels, automobile fuel supersonics jets (beyond 10 ppm retards photosynthesis.	Decomposes ozone. Retards photosynthesis. Causes acid rain.
Oxides of S (SO_2 & SO_3)	Burning of fuels containing sulphur & roasting of sulphide ores (beyond 10ppm] inhibit photosynthesis	Causes respiratory tract deseases. Bleaches green leaves Forma acid rain Prevents photosyntheis Lung cancer, bronchitis, Emphysema.

19. Ozone plays a vital role in converting NO to HNO_3.
20. Normal unpolluted rainwater is weakly acidic 5.6. This is due to dissolved CO_2 gas forming H_2CO_3.
21. pH of acid rain is 4.0 it 1000 times more acidic than pure water. Oxides of nitrogen & sulphur react in the water to form H_2SO_4 & HNO_3.
22. Due to acid rain, monuments made of limestone & marble are decomposed.
23. Carcinogenic benzo pyrene is released from emission tubes of diesel engines.

24. Hydrocarbons are oxidized & converted to free-radicals which are harmful to plants as they damage tissues & cellulose present in it.
25. Peroxy acetyl nitrate (PAN) – hydrocarbon damages plants (0.02 to 0.05 ppm is enough to cause damage).
26. Peroxy benzoyl nitrate (PBN) causes irritation to eyes. It undergoes photochemical oxidation & mixed with ozone & oxides of nitrogen to form smog. [smoke +fog] = smog.
27. In winter smog is formed due to combustion of coal & petroleum.
28. In summer it is formed due to high intensity of solar radiation. Summer smog is oxidizing.
29. Winter smog is reducing in character, it has SO_2 & carbon particles
30. Global warming:- CO_2 & water vapour absorb infrared radiation coming to the earth & partly reflect it back to the earths surface. Due to this the surface of the earth is heated up & is warm even at night.
31. This phenomenon of heating up of the earth surface is called green house effect.
32. Due to de-forestatation the amount of CO_2 is increased in the atmosphere.
33. A 50% increase in CO_2 level increases the surface temp by 3^0c this is called global warming. As a result of global warming the ice caps of the polar region melt here by the level of sea water increases & floods are caused. Agricultural sector gets affected due to fast evaporation of surface water.
34. Green houses gases are CFC 17%; O_3-18%, NO-4%, water vapour 2% mostly CO_2.
35. Ozone is helpful in stratosphere but harmful in troposphere.
36. Ozone layer is affected by the following chemical compounds.
    a) chlorofluoro carbons. b) nitric oxide (NO) c) chlorine.
    The depletion of O_3 layer by 1% increase the incidence of U.V rays on earth by 1-3%. U.V rays produce skin diseases, cancer and cataract of eyes.
37. **Water pollution:** Causes of water pollution.
    i) Waste substance releases from industries.
    ii) Domestic waste due to over – population.
    iii) Green revolution which cause increases in use of pesticides fungicides, insecticides, fertilizers etc.
    iv) Blue revolution in fish ponds.
38. Changes in water due to pollution :-
    Colour & salinity changes, bad odour of water from river, pond, lake, decrease in growth of fish & uncontrolled growth of weeds.
39. Effect of water pollution:- Unfit for drinking, contaminated diseases like cholera, jaundice, typhoid, diarrhea, aquatic life gets destroyed.
    -Decreases in number of tourist visiting the beaches.
40. Bio-amplification:- Polluting substances are carried in food chains & are concentrated (plants – fish –animals) the increases in concentration with the food chain is bio-amplification.
41. Eutrophication:- Polluted water has salts (phosphates) in more quantity this increases nutritive capacity of lakes, it promotes growth of algae (algal bloom) from sediments & cause drying up of ponds. This is called eutrophication.

Water pollutant	Effect
1) Inorganic pollutants eg:- salts, trace elements like Cu, Zn as etc. ii) Metal & complex compounds	Human health & aquatic animals Distinct water system. Algae does not grow. This decreases photosynthesis & increase pollution.
2) Cyanides, H_2S, CO_2, NO_2 & sulphites.	pH of water changes, it becomes toxic,
3) Metals like lead & mercury	Toxic water
4) Fluorides	Bones & teeth get effected (fluorosis) bones become weak, teeth then yellow.
5) Organic pollutants i) Industrial waste ii)Sewage from domestic	Water becomes toxic. Consumes dissolved oxygen.

42. In A.P fluoride is present in drinking water of Nalgonda, Prakasam and certain parts of Guntur districts.
43. On drinking water, containing fluorides. Fluorine reacts with calcium as $Ca + F_2 \rightarrow CaF_2$
44. Fluorides in water can be detected by zirconium – alizarin-s-dye. The colour of dye becomes pale.
    Fluroides < 1ppm – causes dental cavities.
    >3 ppm – causes fluorosis.
45. Water is purified by de-fluoridation, activated carbon method, Nalgonda technique ($CaOCl_2$, lime & alum) etc.
46. Water containing a high concentration of fluoride ions is passed through filters packed with activated carbon which adsorbs the fluoride ions.
47. The exhausted carbon is regenerated by washing with 4% NaOH and 1% H_3PO_4.
48. Water is passed through the filters filled with synthetic resins like Defluoron-I and Defluoron-II to remove fluorine.
49. NEERI institute located at Nagpur discovered an inexpensive method known as the Nalgonda technique.
50. In Nalgonda technique equal quantity of bleaching powder ($CaOCl_2$), lime and alum is used.
51. Sound from traffic, railways, air ports, industries etc., causes sound pollution. Human range of audibility (20-2000 Hz or 60 decibels).
    Sound pollution causes insufficient sleep & leads to physical & mental disorder, high B.P, heart problems, insomnia & irritable temperament.
52. Accidents due to air pollution :-
    i) 45 tons of methyl isocyanate (MIC) released in the air from union carbide factory at Bhopal killed many people.
    ii) Accident in oil refinery at HPCL at vishakapatnam.
    iii) Accident at ONGC well in Konaseema.
    iv) Minimata disease in Japan caused due to diethyl mercury ($Hg(CH_3)_2$) in which people are crippled.
    v) Fire in indoneisa & nearby islands.
53. Ganga water authority cleans ganga water.
54. Oil dispersion in water causes pollution & inhibits photosynthesis.

Fly ash: These are the small ash particles formed by the combustion of fossil fuels which are carried into the air by gases produced during combustion. These enter into the atmosphere from power plants, smelters and mining operations.

Asbestos dust: This originates from industrial units manufacturing asbestos sheets, gaskets, ropes etc. Asbestos insulation and asbestos flooring also add towards asbestos dust in the atmosphere.

Organic Particulars matter: Organic particulates are released in the air when hydro carbons present in coal and oil are oxidized. Among these polycyclic aromatic hydrocarbons (PAH) containing fused benzene. These particles are readily adsorbed on the surface of soot particles present in the atmosphere and therefore are serious health hazards.

Effects of particulars: The effects of particulate pollutants depend upon the size of the particles. The coarser particles of size more than 5 microns are likely to lodge in the nasal passages whereas the smaller ones are more likely to penetrate in to the lungs. The rate of penetration is inversely proportional to the size of the particles. Some of these particles are carcinogens. Continuous inhaling of these small particles are carcinogens. Continuous inhaling of these small particles of long periods of time irritates the lungs and causes 'scarring' or fibrosis of the lung lining. This type of disease is very common in industrial settings and is known as 'pneumoconiosis'.

The suspended particulate matter in the atmosphere reduce visibility and influence weather also. These reduces amount of light rays reaching the surface on the earth and therefore, lowers the temperature of the earth. By blocking the sunlight, they cause drop in earth's temperature by providing condensation nuclei, they contribute to increase to increase fog and rain in cities.

The large concentration of these particles spoil many articles and blackens buildings. These particulates have large surface area hence offer good sites for adsorption of many harmful inorganic and organic matter. Gases like So_2, PAH particles toxic metals (Be, Pd, Cd, Cr, etc.,) are adsorbed on the surface of soot as shown below.

Control of Particulates

Various methods are employed to reduce the presence of carbond and other particulates in the atmosphere.

(i) Effluent gases from the industries are led into a chamber where the velocities of these gases decrease so that dust or droplets get settled. The method is suitable for particles which are settled. This method is suitable for particles which are settled readily and is not for the fine particles which require longer settling time.

(ii) The carbon and other dust particles are removed by electrostatic precipitation. This method is based on the principle that the particles acquire electrical charges when subjected to an electric field. In this method, two electrodes are fitted against the inside walls of the smoke stack. When high voltage (30,000 – 100,000 V) is applied, an electric discharge takes place across the stack. The particles get negatively charged and attracted to the positive electrode located on the inner wall of the stack and settle down there. These can be dislodged from the electrodes by water or vibration of the electrodes and are collected in a resvoir. In this way about 99% of the particulate matter gets removed from the fuel gases.

(iii)The solid, liquid or gaseous particulates can be removed by spraying water from spray chambers.

(iv)Large amount of particles can be collected in cyclone collectors to form mists. Similarly, sulphuric acid mist is produced when So_3 present in the atmosphere comes in contact with moisture. Nitric acid mist produced when oxides of nitrogen (No and No_2) undergo the following reactions.

3. Fumes: These are condensed vapours. For example fumes of metal are very common types of particulars. Some other examples are metallurgical fumes and alkali fumes.

4. Dust: The solid particles suspended in the air are called dust. These particles enter into atmosphere from natural, domestic and industrial sources. These enter the atmosphere by volcanic cruptions, blowing of dust by mining operations, crushing and grinding etc the non-viable dust particulate in the atmosphere consist of ground limestone, sand tailings from floatation pulverized coal, cement, fly ash and silica dust.

These above types of particulars consist of different types of metals and their compounds and organic particulates as discussed below:

(a) Metal particles: These are released into the atmosphere by various metal finishing processes. The common metallic particles in air are lead, mercury, chromium, arsenic, zinc, cadmium, nickel, iron, etc. Although these metallic particles are present in air in very small amounts yet some of these are very dangerous of living organisms. The micro particles of toxic metals and SO_2 gas present in the polluted air get adsorbed on the particles rendering them highly toxic.

(b) Metal oxides: Metal oxides are released into the atmosphere during combustion of fuels containing metallic compounds. For example particulate Fe_3O_4 is produced by the combustion of coal containing iron pyrite as:

$$3FeS_2 + SO_2 \rightarrow Fe_3O_4 + 6SO_2$$

Power plants using vanadium rich residual oil are sources of V_2O_5 pollutants.

(c) Metals salts: Burning of coal gives $CaCo_3$ in ash which is partly converted into CaO. Sulphuric acid droplets in the atmosphere combine with basic pollutants CaO or NH_3 to from salt particulars.

$$H_2SO_4 + CaO \rightarrow CaSO_4 + H_2O$$

$$H_2SO_4 + 2NH_3 \rightarrow (NH_4)_2SO_4$$

(d) Tetraethyl lead: It is a source of lead salts in the atmosphere. Tetracthyl lead is added to gasoline to improve its anti knocking property. It is oxidized to PbO which tends to deposit on spark plugs and values. This is undesirable and is a source of air pollution. To prevent the deposition of PbO, suitable amouns of $C_2H_4Cl_2$ and $C_2H_4Br_2$ are added to the gasoline along with tetracthyl lead. These oxides convert PbO to $PbCl_2$ and Pb Br_2 both of which are volatile and get eliminated through exhaust.

$$2Pb(C_2H_5) + 32O_2 + C_2H_4Br_2 + C_2H_4Cl_2 \rightarrow 20Co_2 + 24H_2O + PbCl_2 + PbBr_2$$

Thus, while the engine of the automobile is saved, the atmosphere gets polluted with dangerous lead salts $PbCl_2$ *and* $PbBr_2$.

(iv) **Non-degradable pollutants:** These include chemicals, mercuric salts, lead compounds, pesticides, etc.

(v) **Natural pollution:** It is caused by Radioactive substances. Volcanic eruptions, forests and mines fires floods etc.

2. **Air Pollution:** The major air pollutants are.

Gaseous air pollutants Co, Co2, So2, No2, No, ammonia, chlorine and HCl Mercaptants.

Particulates:

The small seized solid particles and liquid droplets which range in size from $2\times10^{-10}m$ $(0.0002\mu m)$ to $5\times10^{-4}m$ are collectively called are particulates. These particles are usually individually not visible to the naked eye. However, small particles often collectively from a haze that restricts the visibility. The particulate matter in the range 0.001n to 10 mm range is commonly visible as suspended in the air near the sources of pollution.

These particulates include carbon (soot) metals, metal salts, cement dust, fly ash, hydrocarbons, soil dust. Pulverized coal, asbestos dust, mineral particles etc. Of these, the black smoke released into the air by a diesel engine from trucks, cars or buses is very common form of pollution that we routinely see. The soot particles may deposit on clothing, automobiles, buildings are vegetation. These can blacken sky and act as carriers for other gaseous pollutants. The density of particulates varies from a few hundred per cm^3 in clean air to more than 100,000 per cm^3 in highly polluted air. The pollutants also include liquid aerosols such as sulphuric acid and nitric acid mists.

The particulates in the atmosphere may be viable or non-viable. The viable particulates are the small living small organisms which are dispersed in the atmosphere. These includes bacteria, mould, fungi, algae, etc. Some of these viable particulates cause allergic reaction on human beings. Fungi can also cause plant diseases.

Non-viable particulates: are formed either by the breakdown of large materials or by the condensation of minute particles and droplets.

Types of Particulates and their sources.

There are four types of non-viable particulates in the atmosphere injected into the atmosphere through human activity.

1. **Smoke:** These are very small soot particles. Oil smoke, tobacco smoke and carbon smoke are typical examples of this type of particulate materials. Their common sources are:

(i) Carbon particles are produced in the atmosphere when fossil fuels (coal, fuel oil, natural gas, wood etc) are burnt in an insufficient supply of oxygen.

(ii) Forest fires, coal refuse burning and agriculture burning also produce a lot of smoke.

(iii) Carbon particles are also produced in the atmosphere due to in sufficient burning of fuel in vehicles.

2. **Mists:** These are suspensions of liquid particles in air which arise because of chemical reactions and condensation of vapours in air. For example, mists are produced by the particles of the spray liquids. The common examples are portions of herbicides and insecticides which miss their targets and travel through the air.

Question Bank:

1. Non-harmful level of fluoride in water is
   (1) 2 ppm (2) 5 ppm
   (3) 12 ppm (4) 45 ppm

2. The process of propagation of toxic substances and increase of concentration from lower animals to higher animals are called
   (1) food chain – bioamplification
   (2) bioamplification – eutrophication
   (3) fluorosis – chlorosis
   (4) fluorosis – bioamplification

3. The gas that is used as refrigerant is
   (1) tetrafluoro ethylene
   (2) dichlorodifluoro methane
   (3) trichloro nitromethane
   (4) oxygen diflouride

4. The metal which escapes into atmosphere by fumes of automobiles
   (1) Pb (2) Hg (3) As (4) Cd

5. Eutrophication of lakes is due to
   (1) higher DO value
   (2) over nutrients
   (3) chlorosis
   (4) fluorosis

6. Chemical that is most damaged by acid rain is
   (1) $CaSO_4$ (2) $CaCO_3$
   (3) $CaCl_2$ (4) $Ca(NO_3)_2$

7. The extent of the following green house gas is least
   (1) CO_2 (2) water vapor
   (3) O_3 (4) CRC

8. More toxic water pollutant is
   (1) $CuSO_4$ and $ZnSO_4$
   (2) Hg
   (3) O_2 and N_2
   (4) Na_2CO_3, $NaHCO_3$

9. Ozone hole cannot be created when the following reaction takes place
   (1) $ClO^{\bullet} + NO_2 \rightarrow ClONO_2$
   (2) $Cl^{\bullet}\ O_3 \rightarrow ClO^{\bullet} + O_2$
   (3) $ClO^{\bullet} + O_3 \rightarrow Cl^{\bullet} + 2O_2$
   (4) $CFCl_3 \rightarrow C^{\bullet}FCl_2 + Cl^{\bullet}$

10. In the deflouridation of Nalgonda method *A*, *B* and *C* are added in the order. The *A*, *B* and *C* are

	A	*B*	*C*
(1)	$CaOCl_2$	alum	$Ca(OH)_2$
(2)	$CaOCl_2$	alum	NaOH
(3)	$CaOCl_2$	$Ca(OH)_2$	alum
(4)	$CaOCl_2$	$Ca(OH)_2$	Na_3AlF_6

11. BOD is generally expressed for a time period of
    (1) 5 hours (2) 5 minutes
    (3) 5 months (4) 5 days

12. Regarding eutrophication the wrong statements are
    i) eutrophication lakes are over nutritious
    ii) it decreases the growth and decay of algae
    iii) due to sedimentation the lakes become dry
    iv) it facilitates the luxurious growth of aquatic animals
    (1) ii & iii (2) ii & iv
    (3) i & iii (4) i & ii

13. Phosphates are usually released into water by the use of
    (1) pesticides (2) detergents
    (3) fertilizers (4) perfumes

14. The percentage of water present in the form of ice on earth at poles is

(1) less than 1% (2) about 2%
(3) about 20% (4) about 60%

15. The resin dinuoron-I can remove
(1) Ca^{2+} ions
(2) $Fe^{\surd}$
(3) inorganic pollutants
(4) all of these

16. Identify the correct decreasing order of the following with respect to attitude from atmosphere
(a) troposphere
(b) mesosphere
(c) thermosphere
(1) b, c, a (2) c, b, a
(3) a, b, c (d) a, c, b

17. A substance which grows in quantity due to human activity and adversely effects the environment is called
(1) pollutant (2) contaminent
(3) receptor (4) sink

18. A substance which is not present in nature but released during human activity and adversely effects the environment is called
(1) pollutant (2) contaminent
(3) sink (4) speciation

19. Which one of the following is a contaminent?
(1) SO_2 (2) CO
(3) Pb (4) MiC

20. The amount of oxygen present in water is called
(1) D.O (2) C.O.D
(3) B.O.D (4) T.L.V

21. The amount of oxygen required to oxidize organic substances present in water is called
(1) D.O (2) C.O.D
(3) B.O.D (4) none of these

22. Chemical oxygen demand can be determined by oxidizing the organic substances present in water using
(1) acidified $KMnO_4$
(2) acidified $K_2Cr_2O_7$ solution
(3) chlorine water
(4) bleaching powder

23. The amount of oxygen used by the microorganism present in water for five days is called
(1) C.O.D (2) B.O.D
(3) D.O (4) None

24. The extent of water pollution can be determined by
(1) C.O.D (2) B.O.D
(3) both of these (4) none of these

25. Assertion: Earth's surface temperature is gradually increasing
Reason: CO_2 and H_2O vapour, partly reflects I.R radiation back to earth's surface
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

26. Assertion: Eurtrophic lake becomes over nutritious for algae
Reason: Organic substances from agriculture and industry are thrown into the water resources like ponds and lakes
(1) Both A and R are true and R is the correct explanation of A

(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

27. Assertion: Scientists observed holes in the ozone layer
Reason: Skin diseases cataract cases increased among the people
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

28. Assertion: Greater the COD value of water, greater is its pollution
Reason: COD is the chemical oxygen demand. More COD indicates the presence of more amount of organic waste matter in the polluted water
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

29. Assertion: CO pollution is very high from 9 am to 10 pm in urban areas
Reason: The acceptable level is CO gas is 9 ppm
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

30. Match the following questions:

List –I	List –II
A) Troposphere	I) Eutrophication
B) Mesosphere	II) Acid rain
C) Stratosphere	III) Defluoridation
D) Thermosphere	IV) Green house effect
	V) Bio –amplification

(1) A→I, B→IV, C→V, D→II
(2) A→IV, B→I, C→V, D→II
(3) A→IV, B→I, C→II, D→V
(4) A→I, B→IV, C→II, D→V

31. Match the following:

List –I	List –II
A) Troposphere	I) Non propagation of sound waves
B) Mesosphere	II) Ionisation of gases
C) Stratosphere	III) Non propagation of heat
D) Thermosphere	IV) Maintenance of heat balance
	V) Protection from UV rays

(1) A→IV, B→I, C→V, D→II
(2) A→I, B→IV, C→V, D→II
(3) A→IV, B→I, C→II, D→V
(4) A→I, B→IV, C→II, D→V

KEY

1) **1** 2) **1** 3) **2** 4) **1** 5) **2** 6) **2** 7) **3** 8) **2** 9) **1** 10) **3**
11) **4** 12) **2** 13) **3** 14) **2** 15) **2** 16) **2** 17) **1** 18) **2** 19) **4** 20) **1**
21) **2** 22) **2** 23) **2** 24) **3** 25) **1** 26) **1** 27) **2** 28) **1** 29) **2** 30) **2**
31) **1**

14. ORGANIC CHEMISTRY

Synopsis:

1. According to Berzelius, inorganic compounds are from non - living sources like minerals and organic compounds are from living organisms and these could be produced by 'Vital force'.
2. By heating ammonium cyanate, urea the first organic compound was prepared by Wohler which ruled out 'Vital force theory'.
3. The main natural sources of organic compounds are coal, petroleum, natural gas, animals and plants.
4. Carbon has high catenation capacity and usually tetravalent and forces covalent bonds in organic molecules.
5. The nature of covalent bonding in organic compounds can be described in terms of hybridisation concept, which is sp^3, sp^2 and sp.
6. If all the four valencies of a carbon are satisfied by single bonds, it is called saturated carbon and is in sp^3 hybridisation.
7. If a carbon fonns one pi bond, it is in sp^2 hybridisation and if a carbon forms two pi bonds, it is in 'sp' hybridisation.
8. Methods of purification of solids are : filtration, crystallisation, fractional crystallisation, sublimation, etc.
9. Methods of purification of liquids are: simple distillation, fractional distillation, distillation under reduced pressure, steam distillation, extraction with solvent, etc.
10. The solids which sublime readily on heating are purified by sublimation.
11. Crystallisation is the most common way of purifying organic solids.
12. If two compounds have different solubilities in the same solvent, they are purified by fractional crystallisation.
13. Distillation is the most important method of purifying organic liquids.
14. If the difference in the boiling points of two liquids is more than 40°C, they can be purified by simple distillation, otherwise by fractional distillation.
15. Distillation under reduced pressure is the method used for purification of liquids which decompose at or below their boiling points.
16. Liquids which are immiscible with water, volatile in steam are purified by steam distillation.
17. The process of isolating an organic compound from its aqueous solution by shaking with a suitable solvent is called differential extraction.
18. Chromatography is the most useful and modem technique used to separate mixtures into their components, purify compounds and to test the purity also.
19. Column chromatography and thin layer chromatography are two types of adsorption chromatography which depend upon the principle of differential adsorption.
20. Retardation factor is the ratio between the distance moved by the substance from base line and the distance moved by the solvent from base line.
21. Partition chromatography is based on continuous differential partitioning of components of a mixture between stationary and mobile phases.
22. By heating organic compound with cupn^ oxide, both carbon and hydrogen can be detected together. Carbon is converted to CO, and tested with lime water and

hydrogen is converted to water and tested with anhydrous $CuS0_4$.

23. Lassaigne's test is used for detecting nitrogen, sulphur and halogens. These elements are converted to sodium cyanide, sulphide and halide respectively which are ionic.
24. Lassaigne's extract gives prussian blue colour with ferrous sulphate and concentrated sulphuric acid.
25. Sulphur containing organic compounds give black precipitate with acetic acid and lead acetate solution.
26. Sodium fusion extract gives purple colour with sodium nitroprusside if the organic compound contains sulphur.
27. When sodium fusion extract is acidified with ***l*** dilute nitric acid and treated with silver nitrate solution, precipitate is formed if halogen is present in the organic compound.
28. If phosphorus is present, yellow colouration or precipitate of ammonium phospho- molybdate is formed when organic compound is fused with fusion mixture, boiled with nitric acid and then treated with ammonium molybdate.
29. Oxygen present in the organic compound is detected indirectly.
30. Nitrogen present in the organic compound is estimated by Dumas method or Kjeldahl's method.
31. In Dumas method, nitrogen gas is liberated and in Kjeldahl's method, nitrogen is converted to ammonia.
32. Halogens are estimated by Carius method.

    Percentage of Halogen =

    $$\frac{\text{Atomic mass of X}}{108 + \text{Atomic masss of X}} \times m \times \frac{100}{W}$$

    Where X is Cl, Br or I; m is mass of silver halide formed and W is the mass of the organic substance.
33. Estimation of sulphur is based on the principle of conversion of sulphur into barium sulphate precipitate.
34. Phosphorous present in the organic compound is precipitated as ammonium phosphomolybdate or magnesium pyrophosphate and then estimated.
35. Oxygen is estimated indirectly or by converting into carbondioxide.
36. By using CHN elemental analyser with 1 - 3 mg of sample, elements carbon, hydrogen and nitrogen present in the compound can be determined in a very short time.
37. Organic compounds are mainly classified as acyclic and cyclic componds. Cyclic compounds are classified as homocyclic and heterocyclic. Homocyclic may be alicyclic or aromatic. Aromatic compounds may be benzenoid or non - benzenoid.
38. Heterocyclic compounds may be aromatic or non - aromatic.
39. Acyclic compounds are classified as saturated and unsaturated. Alkenes, alkynes and their derivatives are unsaturated.
40. Furan, pyrrole, thiophene, pyridine, quinoline, etc., are heterocyclic and are aromatic in nature.
41. Planar ring compounds with conjugate double bonds and obeying Huckel's rule [(4n + 2) 7t electrons] are aromatic in nature. Tropone, azulene, etc., are non - benzenoid aromatic compounds.

42. An atom or a group of atoms in an organic substance which is responsible for the characterstic chemical properties is called a fimcional group. Compounds containing the same functional group exhibit similar chemical properties.
43. Organic compounds belonging to the same homologous series exhibit similar chemical properties, regular gradation in physical properties, the successive members differ by - CH_2 - unit and so 14 in molecular weight, can be represented by a general formula and have similar methods of preparation.
44. A given organic compound can be assigned only one name and a given name can have only one molecular structure.
45. According to IUPAC nomenclature, prefix(es) + root word + primary suffix + secondary suffix is the order in naming an organic compound.
46. Primary prefix is meant for alicyclic compounds. Secondary prefix tells us about the substituents or secondary grade functional groups.
47. Primary suffix tells us about saturated or unsaturated compound.
48. Secondary suffix indicates the nature of the functional group.
49. A carbon atom attached to only one carbon atom directly is called primary (1°) carbon, attached to two other carbon atoms is secondary (2°), to three other carbon atoms is tertiary (3°) and to four carbon atoms is quaternary (4°) carbon.
50. Hydrogens attached to primary, secondary and tertiary carbon atoms are termed as primary, secondary and tertiary hydrogen atoms respectively.
51. While naming a carbon compound according to IUPAC nomenclature, longest continuous chain of carbon atoms is to be selected as the parent chain.
52. If some carbon carbon multiple bond is present, the parent chain must include that.
53. It two equally long chains are possible, the chain with maximum number of side chains is selected as parent chain.
54. The numbering of carbon atoms is done in such a way that the substituted carbon atoms have the lowest possible numbers.
55. The functional group bearing carbon should be given the lowest number or lowest sum rule is to be followed.
56. In case of unsaturated hydrocarbons, the carbon atoms involved in the multiple bond should get the lowest possible number.
57. If a number of groups are attached to the parent chain, the name is given to the compound following the increasing order of complexicity of the group (IUPAC) or alphabetical order is followed (CAS). However, CA system is more convenient.
58. The order of decreasing priority for some functional groups is: carboxylic acid > nitrile > aldehyde > ketone > alchol > amine > double bond > triple bond.
59. The chain terminating groups like - COOH, – CHO, – CN, etc., should always get the number ‘ 1 ’ for their carbon atoms irrespective of the above rules.
60. When a benzene ring is attached to an alkane with a functional group, it is considered as substituent, instead of a parent. When no simple base name other than benzene is possible, the positions are numbered so as to give the lowest locant at the first point of difference.
61. Carbon compounds with same molecular formula but different properties are called as isomers and that phenomenon is isomerism.

62. Isomerism is classified mainly into two types: Structural and stereo isomerism.
63. Structural isomerism is due to the difference in the linkage of atoms or groups without any reference to space.
64. Structural isomerism is divided into chain, position, functional group, metamerism and tautomerism.
65. Stereoisomerism is due to the difference in arrangement of atoms or groups in space. Stereo isomers have same molecular formula and also same structural formula. They differ in the spatial arrangement of atoms or groups. It is divided into configurational and conformational isomerism.
66. Chain isomerism arises due to difference in the arrangement of carbon atoms constituting the chain. They differ in the nature of the carbon chain (skeleton).
67. Butane has 2 chain isomers; pentane - 3; hexane -5; heptane - 9; octane - 18; nonane-35; decane - 75 chain isomers. .
68. Position isomerism is due to the difference in the position of a functional group, multiple bond or substituent in the same carbon chain.
69. Compounds having same mo Secular formula but different functional groups are called functional isomers. Alcohols -ethers; aldehydes - ketones; carboxylic acids - esters; cyanides - isocyanides; nitroalkanes - alkyl nitrites; primary, secondary and tertiary amines, alkynes - alkadienes; alkenes - cyclo alkanes; alkynes -cyclo alkenes; etc,, are functional isomers. Functional isomers exhibit different chemical properties along with $_v$ difference in physical properties.
70. Metamerism is due to the presence of different alkyl groups attached to the same bivalent functional group. Metameres differ in the nature of both alkyl groups present on bivalent functional group. Ketones; secondary amines; ethers; etc, show metamerism.
71. The electronic displacements in covalent bonds may occur either due to the presence of some atom or group in the molecule or under the influence of attacking reagent.
72. Inductive effect is the permanent displacement of electrons along a carbon chain when some atom or group of atoms with different electronegativity than carbon is attached to carbon chain.
73. The atom or group which has more power to attract (withdraw) electrons in comparison to hydrogen is said to have -1 effect. The order is : $NO_2 > CN > SO_3H > CHO > CO > COOH > \mathbf{COCl} > COOR > CONH_2 > F > Cl > Br > I > OR > C_6H_5$.
74. The atom or group which has less power to attract the electrons than hydrogen is said to have + I effect. The order is : $(CH_3)C > (CH_3)_2 CH > CH_3 CH_2 > CH_3$.
75. Inductive effect is a permanent effect and it tends to be insignificant beyond the third carbon atom.
76. Acidic nature order of various carboxylic acids and substituted carboxylic acids is:
    i) $HCOOH > CH_3COOH > C_2H_5 COOH > C_3H_7COOH >$
    ii) $FCH_2COOH > ClCH_2COOH > BrCH_2COOH > ICH_2COOH$.
    iii) $CCl_3 COOH > CHCl_2COOH > CH_2ClCOOH > CH_3COOH$
    iv) $CH_3CH_2CHClCOOH > CH_3CHClCH_2COOH > CH_2CHCl_2 CH_2COOH > CH_3CH_2CH_2COOH$

77. Acidic nature order of some other compounds:

i) 

ii) $CH_3OH > CH_3CH_2OH > (CH_3)_2CHOH > (CH_3)_3COH$

78. Relative strength of bases also can be explained by inductive effect

(Due to steric effect, tertiary amines may not be more basic than secondary amines).

98. Electromeric effect is the complete transfer of the shared pair of pi electrons of a multiple bond to one of the atoms in presence of the attacking reagent.

99. Electromeric effect is a temporary effect and comes into play instantaneously at the demand of the attacking reagent.

100. When the transfer of electrons takes place towards the attacking reagent, it is + E effect. Eg. Addition of acids to alkenes.

101. When the transfer of electrons takes place away from the attacking reagent, it is - E effect, eg. Addition of cyanide ion to carbonyl , compounds.

102. When both inductive and electromeric effects t simultaneously operate, usually electromeric effect predominates.

103. Electromeric effect always facilitates the reaction and never inhibits.

104. Mesomeric or resonance effect is a permanent effect involving the transfer of electrons relayed through pi electrons of multiple bonds or a lone pair of electrons and multiple bonds in a chain of carbon atoms in a molecule.

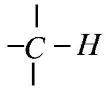
105. Groups with + M effect increase the electron density of the rest of the molecule.
eg. *Cl,* – Br, –1, – NH_2, – NHR, – NR_2,
OH, – OR, – SH, – OCOR, – NHCOR, etc.

106. Groups with - M effect decrease the electron density of the rest of the molecule
eg. – COOH, – COOR, – CHO, – COR,
–CO–, – CN, – NO_2, – SO_3H, etc.

107. Resonance effect is called conjugative effect if it is transmitted through whole of the $_{y}$ conjugated system.

108. The energy of actual structure of the molecule is lower than that of any of the resonance structures. The difference between the real energy of the resonance hybrid structure and the most stable resonance structure is called resonance energy.

109. Resonance explains the stability of aromatic compounds, some unusual bond lengths in some molecules, behaviour of o - and p directing and m - directing groups.

When alkyl groups are attached to an unsaturated system or a benzene nucleus, the sigma electrons present in  $-\overset{|}{\underset{|}{C}}-H$ bond of that alkyl group also involve in conjugation which is called hyperconjugation. It is also called <7 - *it* conjugation or no bond resonance.

110. Greater the number of methyl groups attached to the doubly bonded carbon atoms, greater is the hyper conjugation and greater is its stability.

111. Stability of alkenes, carbocations, alkyl free radicals, orienting effect of alkyl groups in aromatic ring, unexpected bond lengths in some molecules, etc. can be explained by hyperconjugation.

112. Homolytic fission of a covalent bond leads to the formation of neutral species which contain an unpaired electron called free radicals. This is favoured by conditions such as non-polar nature of the bond, high temperature, U.V radiations, presence of peroxides, etc.

113. Heterolytic fission of a covalent bond leads to the formation of charged species. It is favoured by polar nature of the bond, polar solvents, presence of ions due to acid and base catalyst.

114. Carbocation is a group of atoms that contains a carbon atom bearing positive charge and has only six electrons in its valence shell.

115. The positively charged carbon atom in the carbocation is in sp^2 hybridisation. It is trigonal planar in shape.

116. Order of stability of carbocations is:

$$R_3C^+ > R_2\overset{+}{C}H > R\overset{+}{C}H_2 > \overset{+}{C}H_3$$

117. The above stability order can be explained by hyper conjugation and also by +1 effect of alkyl groups.

118. Carbanion is a group of atoms that contains a carbon atom bearing negative charge.

119. The negatively charged carbon atom in the carbanion is in sp^3 hybridisation. Its shape is pyramidal or tetrahedral with one lone pair.

120. The stability order of carbanions is:

$$\bar{C}H_3 > \bar{C}H_3CH_2 > (CH_3)_2\bar{C}H > (CH_3)_3\bar{C}$$

121. If an unpaired electron is present on a carbon atom in a group of atoms which is formed due to homolysis of a covalent bond, it is called alkyl free radical.

122. Alkyl free radicals are planar and the central carbon atom is in sp^2 hybridisation. The unhybridised 2p orbital of the central carbon atom contains the unpaired electron.

123. The stability order of alkyl free radicals is :

$$(CH_3)_3\dot{C} > (CH_3)_2\dot{C}H > CH_3\dot{C}H_2 > \dot{C}H_3$$

The stability order can be explained by hyperconjugation.

124. Carbenes ($\colon CH_2$) are neutral species in which carbon atom has six electrons in the outer shell, out of which two constitute a lone pair and two are shared. In carbenes, carbon is in sp^2 hybridisation.

125. Basing on the nature of the attacking site in the substrate, attacking reagents are classified as nucleophiles and electrophiles.

126. Nucleophiles donate a free electron pair to the electron deficient centre of the organic substrate. Nucleophiles act as Lewis bases.

127. Nucleophiles are either negatively charged or neutral,

eg. Cl^-, Br^-, I^-, $N\bar{H}_2$, RNH^-, R_2N^-, OR^-, $RCOO^-$, R_3C^-, $CH_3COC\bar{H}_2$, OH^-,

CN^-, N_3^-, SH^-, HSO_3^-, etc. are charged nucleophiles and $\ddot{N}H_3$, $R\ddot{N}H_2$, $R_2\ddot{N}H$, $R_3\ddot{N}$, $H_2\ddot{O}:$, $R\underset{\cdot\cdot}{\ddot{O}}R$, $H_2\ddot{S}:$, $R\underset{\cdot\cdot}{\ddot{S}}H$, $R\underset{\cdot\cdot}{\ddot{S}}R$, etc., are neutral nucleophils.

128. Electrophiles are electron deficient and attack the substrate where the electron density is more, act as Lewis acids.

129. Electrophiles are either positively charged or neutral, eg. H^+. Cl^+, Br^+, I^+, NO_2^+, R_3C^+, NH_4^+ , NO^+, $C_6H_5N_2^+$, etc., are charged electrophiles. SO_3, BF_3, $AlCl_3$, $FeCl_3$, $ZnCl_2$, $BeCl_2$, etc., are neutral electrophiles.

130. Ambiphiles are those species which behave like electrophiles and nucleophiles.

eg. $H-\underset{\cdot\cdot}{\ddot{O}}-H$, $R-\underset{\cdot\cdot}{\ddot{O}}-H$, etc.

131. In substitution reactions, an atom or a group attached to a carbon atom in a substrate molecule is replaced by another atom or group.

132. Substitution reactions are further classified as free radical, electrophilic or nucleophilic substitution reactions. Alkanes undergo free radical substitution, aromatic compounds undergo electrophilic substitution and alkyl halides undergo nucleophilic substitution reactions.

133. If the attacking reagent adds on to the substrate molecule without elimination, it is called addition reaction. In addition reactions, a triple bond is converted to double bond and a double bond to single bonds.

134. Addition reactions are also of three types : Electrophilic, nucleophilic or free radical addition reactions. Unsaturated hydrocarbons undergo electrophilic addition, carbonyl compounds undergo nucleophilic addition and addition of HBr to unsymmetrical alkenes in presence of peroxides is free radical addition reaction.

135. In elimination reactions, generally atoms or groups from adjacent carbon atoms in the substrate molecule are removed and a multiple bond is formed. The reverse of addition reaction is known as elimination reaction. In elimination reactions, two sigma bonds are lost and a new pi bond is formed.

136. In the preparation of alkenes; dehydration o: alcohols, dehydrohalogenation of alky' halides, dehalogenation of vicinal dihalides are called p -elimination reactions.

137. In a molecular rearrangement reaction, the product formed is different from that of the expected. Then new compound is actually the structural isomer of the original one.
eg. Fries rearrangement.

138. Aromatic compounds are cyclic, generally containing five, six or seven membered rings, planar with conjugated double bonds. They are resistant towards oxidation and addition reactions under normal conditions.

139. Aromatic compounds are highly stable and obey Huckel rule, i.e., contain $(4n + 2)\ \pi$ electrons where n = 0, 1,2, 3, ...

140. Aromatic compounds containing benzene ring in them are called benzenoid compounds and which do not contain benzene ring are called non-benzenoid compounds.

141. Some ions like cyclopropenyl cation, cyclo heptatrienyl cation, etc. also exhibit aromatic character.

142. Aromatic compounds may be homocyclic or heterocyclic compounds.

143. The carbon-hydrogen ratio in aromatic compounds indicates that they are highly

unsaturated.

144. The two kekule resonance structures of benzene explain 80% properties of benzene and the remaining properties can be explained by the three resonance structures proposed by Dewar.
145. Benzene forms only one monosubstituted derivative which indicates that all the carbon atoms in benzene are identical.
146. Benzene forms three disubstituted products but only one ortho disbustituted product which can be clearly explained by Kekule strucutures.
147. If benzene were a cyclohexatriene, the heat of hydrogenation should be 85.8 Kcals mol^{-1}. But actually it is 49.8 Kcals mol^{-1}. The diffeence of 36 Kcals mol^{-1} is known as resonance energy of benzene.
148. In, benzene all the carbon-carbon bond lengths are equal, (1.39 A^0).
149. Benzene is neither capable of decolourising pink colour of Baeyer's reagent nor orange red colour of bromine in CCl_4,
150. The resonance hybrid structure of benzene has six delocalised Pi electrons.
151. All the carbon atoms in benzene are sp^2 hybridised.
152. The Pi electrons are distributed over all the carbon atoms which is called delocalization. Due to delocaization of Pi electrons, some energy is released which is called as resonance energy.
153. Benzene is formed by the trimerisation (polymerisation) of acetylene in presence of red hot metallic tube.
154. In laboratory, benzoic acid sodium salt when heated with soda lime, benzene is formed.
155. When benzene sulphonic acid is hydrolysed with super heated steam, benzene is formed. Benzene is insoluble in water, but soluble in organic solvents, highly inflammable and burns with a sooty flame.
156. Benzene and other aromatic compounds mainly undergo electrophilic substitution reactions.
157. Generation of electrophile, formation of carbocation and its stabilisation through resonance and removal of proton are the important steps in electrophilic substitution reactions of aromatic compounds.
158. Important electrophilic substitution reactions are nitration, halogenation, sulphonation, Friedel-Crafts alkylation and acylation.
159. in halogenation, halonium ion (X^+);
     in nitration, nitronium ion $(\overset{+}{N}O_2)$;
     in sulphonation, sulphurtrioxide (SO_3);
     in alkylation, alkyl carbocation (R^+) and
     in acylation, acyl carbocation $(\overset{+}{C}OR)$ are the electrophiles.
160. With excess chlorine, all the hydrogen atoms of benzene are substituted by chlorine to form hexachlorobenzene (C_6Cl_6).
161. Under special conditions, benzene undergoes addition reactions also.
162. At high temperature and pressure, benzene adds three moles of hydrogen in presence of nickel to form cyclohexane.
163. In presence of sunlight, benzene adds three moles of chlorine to form benzene hexachloride (BHC) or Gammaxene or Lindane or 666 or hexachlorocyclohexane.
164. Benzene adds three moles of ozone to form unstable benzene triozonide and the final ozonolysis product of benzene is glyoxal.
165. Groups like $-NH_2$; $-NHR$; $-OH$; $-OR$; $-NHCOCH_3$; $-R$; etc., are ortho and para

directing and activating groups.

166. Groups like $-NO_2$; –CHO ; –COR ; –COOH; —COOR; $-SO_3H$; –CN ; etc., are meta directing and deactivating groups.

167. Halogens directly attached to benzene ring, though they are ortho and para directing, they slightly deactivate benzene ring towards further electrophilic substitution reactions.

168. Benzene and several polynuclear hydrocarbons are carcinogenic.

169. Some polynuclear hydrocarbons are formed due to incomplete combustion of organic materials like tobacco, coal and petroleum.

170. Some important carcinogenic polynuclear hydrocarbons are : 1,2–benzanthracene, 3–methylcholanthrene, 1,2–benzpyrene, 1,2,5,6–dibenzanthracene.

171. Benzene is used as a solvent for fats and resins, motor fuel, dry cleaning, manufacture of phenol, styrene, insecticides etc.

Question Bank – I

1. The compound which has one isopropyl group is
(1) 2, 2, 3, 3 – Tetramethyl pentane
(2) 2, 2, 3- trimethyl pentane
(3) 2, 2 - dimethyl pentane
(4) 2 - methyl pentane

2. Two members of a homologous series have
(1) Different general formula.
(2) Different molecular weight.
(3) Different methods of preparation.
(4) Different chemical properties.

3. The compounds in which carbon forms bonds both sp3 and sp2 hybridised orbitals
(1) n – Butane 2. Isobutane
(3) 2 – Butene (4) Butadiene.

4. The IUPAC name of $CH_3 - CH_2 - CH = CH - COOH$ is
(1) pentenoic acid
(2) 2- pentenoic acid
(3) 1- pentanoic acid
(4) 1- pentan -2 – oic acid

5. The following alkynes will not be able to show acidic character
(1) 1-butyne (2) 2-butyne
(3) propyne (4) ethyne

6. Westron vapours when passed over heated $BaCl_2$ gives a compound *Y*. *Y* is a solvent in dry cleaning and extraction of oils from seeds. Then *Y* is
(1) acrylonitrile (2) westrosol
(3) Toluene (4) cresol

7. When acetylene is passed through a solution of $K_2Cr_2O_7 + H_2SO_4$ a compound *X* is formed. Then the functional isomer of *X* is
(1) ehtanoic acid
(2) methyl acetate
(3) acetone
(4) methyl formate

8. Which of these will not react with acetylene?
(1) NaOH
(2) Ammonical $AgNo_3$
(3) Na
(4) HCl

9. A gas decolourises alkaline $KMnO_4$ solution but does not give precipitate with $AgNO_3$. It is
(1) CH_4 (2) C_2H_4
(3) C_2H_2 (4) C_2H_6

10. Ethylene and acetylene can be distinguished by
(1) Baeyer's reagent
(2) Schiff's reagent
(3) Na metal
(4) Schiff's base

11. Among the following a red coloured precipitate is
(1) Na_2C_2 (2) Ag_2C_2
(3) Cu_2C_2 (4) $NaHC_2$

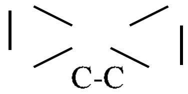
12. Acetylene can be prepared from
(1) Potassium fumarate
(2) Calcium carbide
(3) Ethylene bromide
(4) All of these

13. The hydrogen attached to the following has acidic character

(1)

-C——C-

2) -C≡C-

(3)

C-C

(4) any one

14. In which of the following compound the % of Hydrogen is maximum?
(1) C_2H_2 (2) C_2H_4
(3) C_4H_6 (4) CH_4

15. Acetylene molecule is a neutral compound but its Hydrogen atoms are
(1) acidic (2) basic
(3) neutral (4) amphoteric

16. Benzene was discovered by
(1) Mitscherlich (2) Kekule
(3) Dewar (4) Faraday

17. The resonance energy of benzene is
(1) $1.54 kJmol^{-1}$
(2) $1.39 kJmol^{-1}$
(3) $150.48 kJmol^{-1}$
(4) $140 kJmol^{-1}$

18. Benzene is
(1) Linear (2) Planar
(3) tetrahedral (4) pyramidal

19. Addition of Cl_2 to benzene in the presence of anhydrous $AlCl_3$ is an example of
(1) addition
(2) elimination
(3) substitution
(4) polymerization

20. Gammaxene is
(1) DDT
(2) benzene hexachloride
(3) chloral
(4) hexa chloro ethane

21. Aromatic compounds undergo
(1) nuclephilic addition reaction
(2) electrophilic addition reaction
(3) electrophilic substitution reactions
(4) None

22. Ozone cannot undergo addition with
(1) C_2H_2 (2) C_2H_4
(3) C_6H_6 (4) C_2H_6

23. Which is not used for the preparation of acetylene by Kolbe's electrolysis?
(1) Concentration solution of potassium fumarate
(2) Concentrated solution of potassium maleate
(3) Concentrated solution of sodium succinate
(4) None

24. Gem dihalide on treatment with alcoholic KOH gives

(1) C_2H_2 (2) C_2H_4
(3) C_2H_6 (4) CH_4

25. Acetylene hydrocarbons are acidic because

(1) sigma electron density of C-H bond in acetylene is nearer a carbon which has 50% S-character
(2) acetylene has only one Hydrogen atom at each carbon atom
(3) acetylene contains least number of Hydrogen atoms among the possible hydrocarbons
(4) acetylene belongs to the class of alkynes with formula C_nH_{2n-2}

26. Anhydrous $AlCl_3$ is used in the Fridel-Craft's reaction because it is

(1) Electron rich
(2) Soluble in ether
(3) Covalent
(4) Electron deficient

27. Reaction between benzene and chlorine in presence of halogen carrier like $FeCl_3$ is a good example for

(1) Electrophilic substitution
(2) Nucleophilic substitution
(3) Radical substitution
(4) Electrophilic addition

28. Huckel rule is related to

(1) π electrons
(2) σ electrons
(3) single electrons
(4) number of rings

29. The enolic form of acetone contains

(1) 9σ bonds, 1π bond and 2 lone pairs
(2) 8σ bonds, 2π bonds and 2 lone pairs
(3) 10σ bonds, 1π bond and 1 lone pair
(4) 9σ bonds, 2π bonds and 1 lone pair

30. The most stab le carbonium ion is

(1) $C_2H^+_5$ (2) $(CH_3)_2C^+H$
(3) $CH_3CH_2C^+H_2$ (4) C^+H_3

31. Heterolysis of carbon-chlorine bond produces

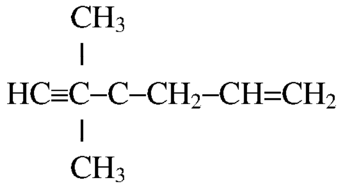
(1) Two free radicals
(2) Two carbonium ions
(3) Two carbanions
(4) One cation and one anion

32. The following are some statements about homologous series

(a) The general formula of a homologous series in C_nH_{2n+2}
(b) The difference between two consecutive members is $-CH_2-$
(c) The are prepared by similar methods

(1) All are correct
(2) Only (a) and (b) are correct
(3) Only (b) is correct
(4) Only (b) and (c) are correct

33. The IUPAC name



$$HC\equiv C-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}-CH_2-CH=CH_2$$

(1) 3, 3 – dimethyl hex –1 –yn –5 –ene

(2) 3, 3 – dimethyl hex –1–en –5– yne
(3) 3, 3 – dimethyl hept –1– en –5– yne
(4) 4, 4– dimethyl hex –1 –en–5–yne

34. Which type of isomerism is shown by pentanone?
(1) Chain isomerism
(2) Position isomerism
(3) Functional isomerism
(4) All of these

35. In the preparation of hydrocarbons by Kolbe's electrolysis, the process involved is
(1) Hydrogenation
(2) decarboxylation
(3) dehydration
(4) dehydrohalogenation

36. A closed vessel contained Beryilium carbide, steam at 1000^0C is passed into the vessel and it is continuously heated. Then the out coming gas mainly contains
(1) CO (2) H_2
(3) CO_2 (4) CO+H_2

37. To prepare a pure sample of n–hexane, using sodium metal as one reactant, the other reactant will be
(1) C_2H_5Cl and C_4H_9Cl
(2) CH_3Br and $C_5H_{11}Br$
(3) C_3H_7Br
(4) C_2H_5Br and C_4H_9Br

38. The bonds present in decane molecule are
(1) 9 C–C bonds and 22 C–H bonds
(2) 7 C–C bonds and 20 C–H bonds
(3) 10 C–C bonds and 5 C–H bonds
(4) 5 C–C bonds and 15 C–H bonds

39. Which of the following reactions will yield 2, 2– dibromopropane?
(1) $CH\equiv CH + 2HBr$
(2) $CH_3–C\equiv CH + 2HBr$
(3) $CH_3CH=CH_2 + HBr$
(4) $CH_3CH=CHBr + HBr$

40. The correct IUPAC name of the compound OHC– CO– COOH is
(1) 2– aldo – 2 – keto methanoic acid
(2) 1 – hydroxy propanoic acid
(3) 3 – formyl – 2 – oxo propanoic acid
(4) 3 – aldo – 2– oxo propanoic acid

41. The sequence of reactions involved in the conversion of acetylene to Toluene are
(1) Trimerisation and nitration
(2) Trimerisation and Polymerisation
(3) Polymerisation and acylation
(4) Trimerisation and alklation

42. An aromatic compound should satisfy the Huckel's rule of $(4n+2)\pi$ electrons. Here n represents
(1) number of rings
(2) number of carbon atoms
(3) zero or integer or fraction
(4) whole number

43. The IUPAC name of CHO – CHO
(1) Glyoxal (2) Dialdehyde
(3) Ethanedial (4) Ethane diol

44. Two organic compounds *A* and *B* both containing only carbon and

hydrogen, on quantitative analysis gave the same percentage composition by weight: *A* decolourises bromine water but *B* does not decolourise bromine water. Then *A* and *B* are

A *B*
1) C_2H_2 C_6H_6
2) C_6H_6 C_2H_2
3) C_2H_4 C_2H_6
4) C_2H_2 C_2H_6

45. A hydrocarbon of formula C_6H_6 decolourises bromine water. It also gives precipitate with ammonical $AgNO_3$ solution. The hydrocarbon is
(1) 1, 3, 5- cyclohexatriene
(2) 1, 5- hexadiyne
(3) 2, 4- hexadiyne
(4) None of them

46. $C_6H_6 \xrightarrow{H_2/Ni} C_6H_{12}$, in this reaction the hybridisation of benzene changes from sp^2 to
(1) sp^2 (2) sp^3
(3) sp (4) sp^3d

47. $CaC_2 \xrightarrow{H_2O} (A) \xrightarrow[\Delta]{\text{Fe tube}} (B) \xrightarrow[\text{high P}]{H_2/Ni} (C)$
Here *C* is
(1) benzene (2) cyclohexane
(3) cyclohexene (4) hexane

48. Nitration mixture is
(1) con HNO_3 + con H_2SO_4 (1:1 by weight ratio)
(2) con HNO_3 + con. H_2SO_4 (1:1 by volume ratio)
(3) con. HNO_2 + con HNO_3 (1:1 by volume ratio)
(4) con HNO_3 + con HCl (1:3 by weight ratio)

49. Aromatic compound contains
(1) $4n + 2\pi$ electrons
(2) $(4n + 2)\pi$ electrons
(3) $(2n + 2)\pi$ electrons
(4) $(4n + 1)\pi$ electrons

50. Assertion: The IUPAC name of the compound $CH_3–CH\ (Cl)\ –CH_2–CH_2–CH_3$ is 2– chloro pentane but not 4– chloro pentane.

Reason: The substituent must be attached to the lowest numbered carbon atom.
(1) Both A and R are true and R is the correct explanation of A
(2) Both A and R are true and R is not the correct explanation of A
(3) A is true but R is false
(4) A is false but R is true

KEY

1) **4**	2) **2**	3) **3**	4) **2**	5) **2**	6) **2**	7) **4**	8) **1**	9) **2**	10) **3**
11) **1**	12) **4**	13) **2**	14) **4**	15) **1**	16) **4**	17) **3**	18) **2**	19) **3**	20) **2**
21) **3**	22) **4**	23) **3**	24) **1**	25) **1**	26) **4**	27) **1**	28) **1**	29) **1**	30) **2**
31) **4**	32) **4**	33) **4**	34) **4**	35) **2**	36) **4**	37) **3**	38) **1**	39) **2**	40) **3**
41) **4**	42) **4**	43) **3**	44) **1**	45) **2**	46) **2**	47) **2**	48) **2**	49) **2**	50) **1**

Question Bank – II

1. Chromatography technique was discovered by
   1) Tswett 2) Schonbein 3) Claisen 4) Perkin
2. Mixture of glucose and water can be separated by
   1) Distillation 2) Crystallisation
   3) Steam distillation 4) Fractional distillation
3. In adsorption chromatography mobile phase will be
   1) only solid 2) only liquid 3) only gas 4) iquid as well as gas
4. Which of the following are correct for purification by steam distillation.
   1) Impurities must be non-volatile
   2) The liquid must be completely immiscible with water
   3) The vapour pressure of liqust should be high
   4) All are correct
5. Two substances when separated on the basis of partition co-efficient between two liquid phase, the technique is known as
   1) Column chromatography 2) Paper chromatography
   3) Fraction distillation 4) Fractional crystallization
6. In column chromatography stationary phase is
   1) Only solid 2) Only liquid 3) Only gas 4) Any of those
7. Paper chromatography is
   1) Adsorption chromatography 2) Partition chromatography
   3) Ion exchange chromatography 4) None of these
8. Separation of components present in mixture by column chromatography is due to
   1) Relative adsorption 2) Relative adsorption
   3) Solubilities 4) Relative adsorption solubilities

KEY

1) **1** 2) **1** 3) **1** 4) **4** 5) **2** 6) **1** 7) **2** 8) **1**

Question Bank – III

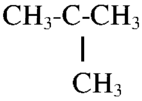
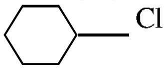
1. Presence of nitrogen in organic compound is tested as
   (1) Nitrogen gas (2) NH_3 (3) No (4) CN^-
2. In lassaigne's method organic compound is fused with
   (1) Sodium metal (2) Zn dust
   (3) Sodium Carbonate and Zn dust (4) Calcium metal
3. Sodium extract gives blood colour when treated with $FeCl_3$ formation of blood red colour confirms in the presence of
   (1) Only nitrogen (2) Only Sulphur (3) Only halogen (4) Both nitrogen & sulphur
4. In which of the following compounds % of nitrogen cannot be estimated accurately by kjclddh's method.
   (1) Urea (2) Phenyhydrazine (3) Nitro benzene (4) Gaunidine

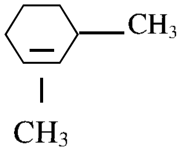
5. Halogen can be estimated by
   (1) Duma's method (2) Carius method (3) Leibig's method (4) All of these
6. 365 of an organic compound containing nitrogen gave 56 ml of nitrogen at S.T.P. The percentage of nitrogen in the given compounds.
   (1) 19.18 (2) 38.36 (3) 9.18 (4) 29.18

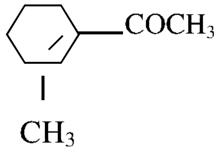
KEY

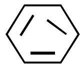
1) **4** 2) **1** 3) **4** 4) **3** 5) **2** 6) **1**

Question Bank – IV

1. Which of the following has high nucleophilicity.
   (1) $F-$ (2) $OH-$ (3) CH_3- (4) NH_2
2. +R power of given group.
   (a) $-O^{(+)}$ (b) $-NH_2$ (c) -OH (d) $-NHCOCH_3$
   in decreasing order is
   (1) 1>2>3>4 (2) 4>3>2>1 (3) 1>3>2>4 (4) 1>4>3>2
3. Which of the following statement are correct for nucleophile
   (1) All negatively charged species are nucleophile
   (2) Nucleophiles are lews bases
   (3) Alkenes, alkynes, benzene and pyrole are nucleophiles
   (4) All are correct.
4. The stability of given free radical in decreasing order is
   (a) $CH_3 - CH_2$ (b) $CH_3\ CH\ CH_3$ (c)  $\underset{\underset{CH_3}{|}}{CH_3\text{-}C\text{-}CH_3}$ (d) CH_3
5. In which of the following molecules all the effects namely inductive, mesomeric, hyper conjugation operate
   (a) 

   (b) 

   (c) 

   (d) 
6. Which of the following pair is correctly reached.
   (1) Carbocation : Electrophile (2) Free radical : Paramagnetic
   (3) Carbane : Incomplete octet (4) Carbonion : Incomplete actet
7. Resonance is due to
   (1) Delocalization of sigma electrons (2) Migration of H atoms
   (3) Migration of Proton (4) Delocalisation of Π electrons
8. Which of the following pair is carbocation is most stable.
   (1) $CH_3 - CH_2 - CH_2$ (2) $CH_3 - CH_2 - CH_2 - CH_2$
   (3) $CH_3 - CH - CH_3$ (4) $CH_3 - CH_2 - CH - CH_3$

KEY

1) **2** 2) **1** 3) **1** 4) **4** 5) **3** 6) **4** 7) **4** 8) **4**