

(CLASS -XI)

FINAL
APPROVED

XI Chemistry Chapter Notes

▪

•

..

Chapter notes

Class: XI

Chapter 1: Some Basic Concepts of Chemistry

Top concepts

1. The SI system has seven base units which pertain to the 7 fundamental scientific quantities

Base Physical Quantity	Symbol for quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	Kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

2. The unit is written on the right of the quantity with a space between them.
3. The SI system allows the use of prefixes to indicate the multiples or submultiples of a unit.

Multiple	Prefix	Symbol
10^{-1}	deci	d

10	deca	da
10^2	hecto	h
10^3	kilo	k
10^6	mega	M

4. To indicate very small numbers, we use negative exponents.
5. To indicate large numbers, we use positive exponents.
6. Scientific notation is the proper representation of a number in exponential form.
7. Precision indicates how closely repeated measurements match each other.
8. Accuracy indicates how closely a measurement matches the correct or expected value.
9. A result is valid only if it is both accurate and precise.
10. Significant figures are meaningful digits which are known with certainty.
11. There are certain rules for determining the number of significant figures:
 - i) All non-zero digits are significant
 - ii) Zeros preceding the first non-zero digit are not significant
 - iii) Zeros between two non-zero digits are significant.

iv) Zeros at the end or right of the number are significant provided they are on the right side of the decimal point. But, if otherwise, the zeros are not significant.

- 14.** During addition and subtraction, the result cannot have more digits to the right of the decimal point than either of the original numbers.
- 15.** In multiplication and division with significant figures, the answer cannot have more significant figures than either of the original numbers.
- 16.** There are 5 basic laws of chemical combinations that govern every reaction: Law of conservation of mass, law of definite proportions, law of multiple proportions, Gay Lussac's law of gaseous volumes, and lastly, Avogadro law.
- 17. Law of Conservation of Mass:** Antoine Lavoisier established the Law of Conservation of Mass. It states that matter can neither be created nor destroyed. In other words, we can say that during any physical or chemical change, the total mass of reactants is equal to the total mass of products.
- 18. Law of definite proportions:** Joseph Proust showed that a given compound always contains exactly the same proportion of elements by weight.
- 19. Law of multiple proportions:** Dalton proposed the law of multiple proportions. According to this law if two elements can combine to form more than one compound, the mass of one element that combines with the fixed mass of the other element is in the ratio of small whole numbers

- 20. Gay Lussac's Law of gaseous volumes:** When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all the gases are at same temperature and pressure.
- 21. Avogadro law:** At the same temperature and pressure, equal volumes of gases contain equal number of molecules.
- 22. Dalton's atomic theory:** In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following :
- Matter consists of indivisible atoms.
 - All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
 - Compounds are formed when atoms of different elements combine in a fixed ratio.
 - Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.
- 23.** Dalton's theory could explain the laws of chemical combination.
- 24.** The number 6.022×10^{23} is called Avogadro's constant or Avogadro's number.
- 25.** A mole is a collection of 6.022×10^{23} particles.
- 26.** One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C
- 27.** The mass of one mole of a substance in grams is called its molar mass.
- 28.** The molar mass in grams is numerically equal to the atomic/molecular/formula mass in u.(u is the unified mass)

- 29.** Molarity is the number of moles of solute in per liter of solution. Unit is moles per liter.
- 30.** Molality is the number of solute present in 1kg of solvent.
- 31.** Atomic Mass: Average relative mass of an atom of an element as compared with the mass of a carbon atom taken as 12 amu
- 32.** Atomic mass expressed in grams is called gram atomic mass
- 33.** Molecular Mass: Sum of the atomic masses of elements present in a molecule
- 34.** Molecular mass expressed in grams is called gram molecular mass
- 35.** Formula Mass: Sum of atomic masses of all atoms in a formula unit of the compound
- 36.** Following relations given below can be summarized
- One mole of atoms = 6.022×10^{23} atoms = Gram atomic mass of element
 - One mole of molecules = 6.022×10^{23} molecules = Gram molecular mass of substance
- 37.** An empirical formula represents the simplest whole number ratio of various atoms present in a compound.
- 38.** Molecular formula shows the exact number of different types of atoms present in a molecule of a compound.
- 39.** If the mass per cent of various elements present in a compound is known, its empirical formula can be determined.
- 40.** Molecular formula = n (Empirical formula), where n is a simple number and may have values 1, 2, 3...
- 41.** Following steps should be followed to determine empirical formula of the compound
- Step 1: Conversion of mass per cent of various elements into grams.
 - Step 2: Convert mass obtained in step1 into number of moles

- Step 3: Divide the mole value obtained in step 2 by the smallest mole value (out of the mole value of various elements calculated)
- Step 4: In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.
- Step 5: Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

- 42.** Anything that has mass and occupies space is called matter
- 43.** Matter can either be a mixture or be a pure substance
- 44.** Homogenous mixtures are those whose components completely mix with each other to make a uniform composition
- 45.** Heterogeneous mixtures are not uniform, and their components are separable through physical methods
- 46.** Pure substances can be elements or compounds
- 47.** An element consists of only one type of particles
- 48.** Two or more atoms of different elements combine to form a molecule of a compound
- 49.** The constituents of a compound can be separated only by chemical methods.
- 50.** A compound has properties different from its constituent elements
- 51.** Isotopes are elements with same atomic number but different mass number.
- 52.** Atomic mass is denoted by "u" – unified mass.
- 53.** One mole is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of the ^{12}C isotope
- 54.** The mass of one mole of a substance in grams is called its molar mass
- 55.** Out of various reactants in a reaction, a reactant that is completely consumed in a chemical reaction is called limiting reagent

- 56.** Stoichiometry gives a quantitative relation between reactant and product in a reaction. It also helps us in identifying limiting reagents

Top Formulae

Mass % of an element

1.
$$= \frac{\text{Mass of that element in the compound}}{\text{Molar mass of compound}} \times 100$$
2. Mass per cent =
$$\frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$
3. Mole Fraction =
$$\frac{\text{No. of mole of a particular component}}{\text{Total No. of moles of solution}}$$
4. Molarity =
$$\frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$
5. Molality =
$$\frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$
6. Moles of an element =
$$\frac{\text{Mass of element}}{\text{Atomic mass}}$$
7. Mass of one atom =
$$\frac{\text{Atomic mass}}{6.022 \times 10^{23}}$$
8. Moles of a compound =
$$\frac{\text{Mass of compound}}{\text{Molecular mass}}$$
9. Mass of one molecule =
$$\frac{\text{Molecular mass}}{6.022 \times 10^{23}}$$

Points To Remember
Class: XI
Ch 2: Structure O Atom

Top Concepts

1. Atomic theory of matter was proposed by John Dalton
2. Electrons were discovered by Michael Faraday.
3. Electrons were discovered using cathode ray discharge tube experiment.
4. Cathode ray discharge tube experiment: A cathode ray discharge tube made of glass is taken with two electrodes. At very low pressure and high voltage, current starts flowing through a stream of particles moving in the tube from cathode to anode. These rays were called cathode rays. When a perforated anode was taken, the cathode rays struck the other end of the glass tube at the fluorescent coating and a bright spot on the coating was developed
Results:
 - a. Cathode rays consist of negatively charged electrons.
 - b. Cathode rays themselves are not visible but their behavior can be observed with help of fluorescent or phosphorescent materials.
 - c. In absence of electrical or magnetic field cathode rays travel in straight lines
 - d. In presence of electrical or magnetic field, behaviour of cathode rays is similar to that shown by electrons
 - e. The characteristics of the cathode rays do not depend upon the material of the electrodes and the nature of the gas present in the cathode ray tube.
5. Charge to mass ratio of an electron was determined by Thomson. The charge to mass ratio of an electron as $1.758820 \times 10^{11} \text{ C kg}^{-1}$
6. Charge on an electron was determined by R A Millikan by using an oil drop experiment. The value of the charge on an electron is $-1.6 \times 10^{-19} \text{ C}$.
7. The mass on an electron was determined by combining the results of Thomson's experiment and Millikan's oil drop experiment. The mass of an electron was determined to be $9.1094 \times 10^{-31} \text{ kg}$.
8. Discovery of protons and canal rays: Modified cathode ray tube experiment was carried out which led to the discovery of protons.

9. Canal rays are positively charged particles called protons

10. Characteristics of positively charged particles:

- a. Charge to mass ratio of particles depends on gas from which these originate
- b. The positively charged particles depend upon the nature of gas present in the cathode ray discharge tube
- c. Some of the positively charged particles carry a multiple of fundamental of electrical charge.
- d. Behaviour of positively charged particles in electrical or magnetic field is opposite to that observed for cathode rays

11. Neutrons were discovered by James Chadwick by bombarding a thin sheet of beryllium by α - particles. They are electrically neutral particles having a mass slightly greater than that of the protons.

12. Thomson model of an atom: This model proposed that atom is considered as a uniform positively charged sphere and electrons are embedded in it.

13. An important feature of Thomson model of an atom was that mass of atom is considered to be evenly spread over the atom.

14. Thomson model of atom is also called as Plum pudding, raisin pudding or watermelon model

15. Thomson model of atom was discarded because it could not explain certain experimental results like the scattering of α - particles by thin metal foils

16. Observations from α - particles scattering experiment by Rutherford:

- a. Most of the α - particles (nearly 99 %) passed through gold foil undeflected
- b. A small fraction of α - particles got deflected through small angles
- c. Very few α - particles did not pass through foil but suffered large deflection nearly 180°

17. Observations Rutherford drew from α - particles scattering experiment:

- a. Since most of the α - particles passed through foil undeflected, it means most of the space in atom is empty
- b. Since some of the α - particles are deflected to certain angles, it means that there is positively mass present in atom
- c. Since only some of the α - particles suffered large deflections, the positively charged mass must be occupying very small space

d. Strong deflections or even bouncing back of α -particles from metal foil were due to direct collision with positively charged mass in atom

18. Rutherford's model of atom: This model explained that atom consists of nucleus which is concentrated in a very small volume. The nucleus comprises of protons and neutrons. The electrons revolve around the nucleus in fixed orbits. Electrons and nucleus are held together by electrostatic forces of attraction.

19. Drawbacks of Rutherford's model of atom:

a. According to Rutherford's model of atom, electrons which are negatively charged particles revolve around the nucleus in fixed orbits. Thus, the electrons undergo acceleration. According to electromagnetic theory of Maxwell, a charged particle undergoing acceleration should emit electromagnetic radiation. Thus, an electron in an orbit should emit radiation. Thus, the orbit should shrink. But this does not happen.

b. The model does not give any information about how electrons are distributed around nucleus and what are energies of these electrons

20. Atomic number (Z): It is equal to the number of protons in an atom. It is also equal to the number of electrons in a neutral atom.

21. Mass number (A): It is equal to the sum of protons and neutrons.

22. Isotopes: These are the atoms of the same element having the same atomic number but different mass number.

23. Isobars: Isobars are the atoms of different elements having the same mass number but different atomic number.

24. Isoelectronic species: These are those species which have the same number of electrons.

25. Electromagnetic radiations: The radiations which are associated with electrical and magnetic fields are called electromagnetic radiations. When an electrically charged particle moves under acceleration, alternating electrical and magnetic fields are produced and transmitted. These fields are transmitted in the form of waves. These waves are called electromagnetic waves or electromagnetic radiations.

26. Properties of electromagnetic radiations:

- a. Oscillating electric and magnetic field are produced by oscillating charged particles. These fields are perpendicular to each other and both are perpendicular to the direction of propagation of the wave.
- b. They do not need a medium to travel. That means they can even travel in vacuum.

27.Characteristics of electromagnetic radiations :

- a. Wavelength: It may be defined as the distance between two neighbouring crests or troughs of wave as shown. It is denoted by λ .
- b. Frequency (ν): It may be defined as the number of waves which pass through a particular point in one second.
- c. Velocity (v): It is defined as the distance travelled by a wave in one second. In vacuum all types of electromagnetic radiations travel with the same velocity. Its value is $3 \times 10^8 \text{ m sec}^{-1}$. It is denoted by v
- d. Wave number: Wave number ($\bar{\nu}$) is defined as the number of wavelengths per unit length.

28.Relationship between velocity, frequency and wavelength

Velocity = frequency x wavelength

$$c = \nu\lambda$$

29.Black body: An ideal body, which emits and absorbs all frequencies, is called a black body. The radiation emitted by such a body is called black body radiation.

30.Planck's quantum theory: Max Planck suggested that atoms and molecules could emit or absorb energy only in discrete quantities and not in a continuous manner. Planck gave the name quantum, meaning 'fixed amount' to the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation.

$$E \propto \nu$$

$$E = h \nu = \frac{hc}{\lambda}$$

Where:

E is the energy of a single quantum

ν is the frequency of the radiation

h is Planck's constant

$$h = 6.626 \times 10^{-34} \text{ Js}$$

31.Quantisation of energy: Energy is always emitted or absorbed as integral multiple of this quantum.

$$E = nh\nu$$

Where $n=1, 2, 3, 4, \dots$

- 32.** Photoelectric effect: The phenomenon of ejection of electrons from the surface of metal when light of suitable frequency strikes it is called photoelectric effect. The ejected electrons are called photoelectrons.
33. Experimental results observed for the experiment of Photoelectric effect observed Hertz:
- When beam of light falls on a metal surface electrons are ejected immediately i.e. there is not time lag between light striking metal surface and ejection of electrons
 - Number of electrons ejected is proportional to intensity or brightness of light
 - Threshold frequency (ν_0): For each metal there is a characteristic minimum frequency below which photoelectric effect is not observed. This is called threshold frequency.
 - If frequency of light is less than the threshold frequency there is no ejection of electrons no matter how long it falls on surface or how high is its intensity.

34. Photoelectric work function (W_0): The minimum energy required to eject electrons is called photoelectric work function.

$$W_0 = h\nu_0$$

35. Energy of the ejected electrons :

$$h(\nu - \nu_0) = \frac{1}{2} m_e v^2$$

36. When a white light is passed through a prism, it splits into a series of coloured bands known as spectrum.
37. Spectrum is of two types: continuous and line spectrum
- The spectrum which consists of all the wavelengths is called continuous spectrum.
 - A spectrum in which only specific wavelengths are present is known as a line spectrum. It has bright lines with dark spaces between them.
38. Electromagnetic spectrum is a continuous spectrum. It consists of a range of electromagnetic radiations arranged in the order of increasing wavelengths or decreasing frequencies. It extends from radio waves to gamma rays.

39. Spectrum is also classified as emission and line spectrum.

- c. Emission spectrum: A substance absorbs energy and moves to a higher energy state. The atoms, molecules or ions that have absorbed radiation are said to be excited. Since the higher energy state is unstable they return to the more stable energy state by emitting the absorbed radiation in various regions of electromagnetic spectrum. The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.
- d. Absorption spectrum is the spectrum obtained when radiation is passed through a sample of material. The sample absorbs radiation of certain wavelengths. The wavelengths which are absorbed are missing and come as dark lines.

40. The study of emission or absorption spectra is referred as spectroscopy.

41. Spectral Lines for atomic hydrogen:

Series	n₁	n₂	Spectral Region
Lyman	1	2, 3, 4, 5 ...	Ultraviolet
Balmer	2	3, 4, 5 ...	Visible
Paschen	3	4, 5 ...	Infrared
Brackett	4	5, 6 ...	Infrared
Pfund	5	6, 7...	Infrared

42. Rydberg equation: It allows the calculation of the wavelengths of all the spectral lines of hydrogen.

$$\bar{\nu} = 109,677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1}$$

43. Bohr's model for hydrogen atom:

- a. An electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy. These paths are called orbits or energy levels. These orbits are arranged concentrically around the nucleus.
- b. As long as an electron remains in a particular orbit, it does not lose or gain energy and its energy remains constant.
- c. When transition occurs between two stationary states that differ in energy, the frequency of the radiation absorbed or emitted can be calculated.

$$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

ν = Frequency of radiation

h = Planck's constant

E_1 = Energy of lower energy state

E_2 = Energy of higher energy state

d. An electron can move only in those orbits for which its angular momentum is an integral multiple of $h/2\pi$

$$m_e v r = n \cdot \frac{h}{2\pi} \quad n = 1, 2, 3, \dots$$

43. Bohr's theory for hydrogen atom:

a. Stationary states for electron are numbered in terms of Principal Quantum numbered as $n=1, 2, 3, \dots$

b. For hydrogen atom: The radii of the stationary states is expressed as $r_n = n^2 a_0$ where $a_0 = 52.9 \text{ pm}$

c. Energy of stationary state

$$E_n = -R_H \left(\frac{1}{n^2} \right)$$

where $R_H = 2.18 \times 10^{-18} \text{ J (Rydberg constant)}$

$n = 1, 2, 3, \dots$

$$E_n = -2.18 \times 10^{-18} \left(\frac{1}{n^2} \right) \text{ J}$$

d. For ions containing only one electron:

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right) \text{ J}$$

where $n = 1, 2, 3, \dots$

$$r_n = \frac{n^2 a_0}{Z} \text{ pm}$$

Where Z is the atomic number

44. Limitations of Bohr's model of atom:

- a. Bohr's model failed to account for the finer details of the hydrogen spectrum. For instance splitting of a line in the spectrum into two closely spaced lines.
- b. Bohr's model was also unable to explain spectrum of atoms containing more than one electron.
- c. Bohr's model was unable to explain Zeeman effect i.e. splitting of spectral line in presence of magnetic effect.
- d. Bohr's model also failed to explain Stark effect i.e. splitting of spectral line in presence of electric field.
- e. Bohr's model could not explain the ability of atoms to form molecules by chemical bonds

45. Dual behavior of matter: de Broglie proposed that matter exhibits dual behavior i.e. matter shows both particle and wave nature.

1. de Broglie's relation:

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

Where:

- λ - Wavelength
- p - Momentum
- v - Velocity
- h - Planck's constant

2. According to de Broglie, every object in motion has a wave character. Wavelengths of macroscopic objects cannot be detected but for microscopic particles it can be detected. This is because for microscopic objects, the mass is less. Since mass and wavelength are inversely proportional to each other, the wavelength will be more. But for macroscopic objects, the mass is large. Therefore, wavelength will be too short to be detected.
3. Heisenberg's uncertainty principle: It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron.

$$\Delta x \cdot \Delta p_x \geq \frac{h}{4\pi}$$

$$\Delta x \cdot \Delta(m v_x) \geq \frac{h}{4\pi}$$

$$\Delta x \cdot \Delta v_x \geq \frac{h}{4\pi m}$$

Where

Δx - Uncertainty in position

Δv_x - Uncertainty in velocity

Δp_x - Uncertainty in momentum

This means that if the position of electron is known, the velocity of electron will be uncertain. On the other hand, if the velocity of electron is known precisely, the position of electron will be uncertain.

4. Heisenberg's uncertainty principle rules out the existence of definite paths or trajectories of electrons and other similar particles
5. Failure of Bohr's model:
 - a. It ignores the dual behavior of matter.
 - b. It contradicts Heisenberg's uncertainty principle.
46. Classical mechanics is based on Newton's laws of motion. It successfully describes the motion of macroscopic particles but fails in the case of microscopic particles.
Reason: Classical mechanics ignores the concept of dual behaviour of matter especially for sub-atomic particles and the Heisenberg's uncertainty principle.
47. Quantum mechanics is a theoretical science that deals with the study of the motions of the microscopic objects that have both observable wave like and particle like properties.
48. When quantum mechanics is applied to macroscopic objects (for which wave like properties are insignificant) the results are the same as those from the classical mechanics.
49. Quantum mechanics is based on a fundamental equation which is called Schrodinger equation.
50. Schrodinger's equation: For a system (such as an atom or a molecule whose energy does not change with time) the Schrödinger equation is written as:

$$\hat{H}\Psi = E\Psi$$

Where:

\hat{H} is the Hamiltonian operator

E is the total energy of the system

Ψ represents the wave function which is the amplitude of the electron

Wave

51. When Schrödinger equation is solved for hydrogen atom, the solution gives the possible energy levels the electron can occupy and the corresponding wave function(s) of the electron associated with each energy level.

Out of the possible values, only certain solutions are permitted. Each permitted solution is highly significant as it corresponds to a definite energy state. Thus, we can say that energy is quantized. That is, it can have only certain specific values.

52. Ψ gives us the amplitude of wave. The value of ψ has no physical significance.
53. Ψ^2 gives us the region in which the probability of finding an electron is maximum. It is called probability density.
54. Orbital: The region of space around the nucleus where the probability of finding an electron is maximum is called an orbital.
55. Quantum numbers: There are a set of four quantum numbers which specify the energy, size, shape and orientation of an orbital. These are:
- Principal quantum number (n)
 - Azimuthal quantum number (l)
 - Magnetic quantum number (m_l)
 - Electron spin quantum number (m_s)
56. Principal quantum number (n): It determines the size and to a large extent the energy of the orbital.

n	1	2	3	4
Shell no.:	K	L	M	N
Total number of orbitals in a shell = n^2	1	4	9	16
Maximum number of electrons = $2n^2$	2	8	18	32

- It can have positive integer values of 1, 2, 3 and so on.
- It also identifies the shell.
- As the value of n increases, the energy also increases. Hence, the electron will be located far away from the nucleus.

57. Azimuthal quantum number (l): Azimuthal quantum number. ' l ' is also known as orbital angular momentum or subsidiary quantum number. It identified the shell and the three dimensional shape of the orbital.
- It also determines the number of subshells or sub levels in a shell. Total number of subshells in a particular shell is equal to the value of n .
 $l = 0, 1, 2... (n-1)$

- Each subshell corresponding to different values of l are represented by different symbols:

Value of l	0	1	2	3
Notation of symbol	s	p	d	f

58. Magnetic quantum number or Magnetic orbital quantum number (m_l): It gives information about the spatial orientation of the orbital with respect to standard set of co-ordinate axis.

For any sub-shell (defined by ' l ' value) $2l+1$ values of m_l are possible.

For each value of l ,

$$m_l = -l, -(l-1), -(l-2)\dots 0, 1\dots (l-2), (l-1), l$$

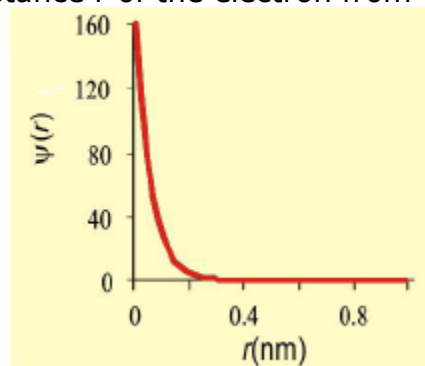
59. Electron spin quantum number (m_s): It refers to orientation of the spin of the electron. It can have two values $+1/2$ and $-1/2$. $+1/2$ identifies the clockwise spin and $-1/2$ identifies the anti-clockwise spin.

60. An orbital is identified by the set of 3 quantum numbers: Principal quantum number, Azimuthal quantum number and magnetic quantum number.

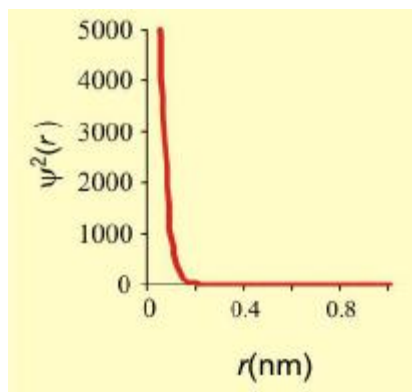
61. An electron is identified by a set of four quantum numbers: Principal quantum number, azimuthal quantum number, magnetic quantum number and spin quantum number.

62. Sub-shell notation: Notation of a sub-shell is written as the Principal quantum number followed by the symbol of the respective sub-shell.

63. Plots of the orbital wave function $\Psi(r)$ and probability density $\Psi^2(r)$ Vs distance r of the electron from the nucleus for 1s orbital:

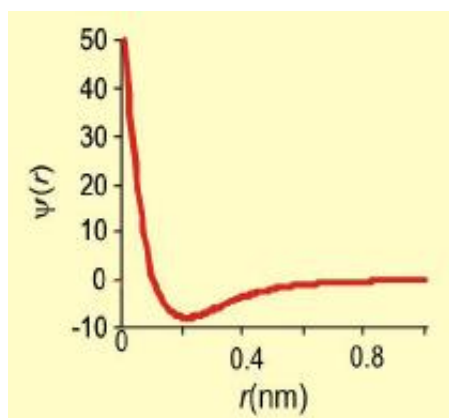


- For 1s orbital the probability density is maximum at the nucleus and it decreases sharply as we move away from it (which is not possible). Hence plot of probability density $\Psi^2(r)$ Vs distance r of the electron from the nucleus was drawn as shown below.

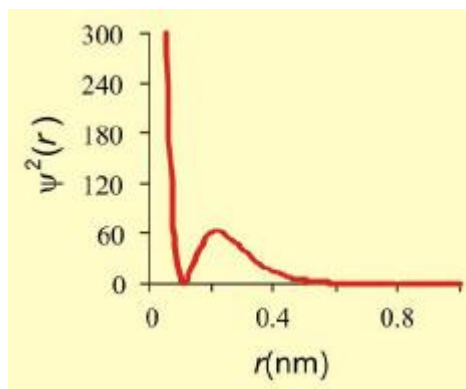


- The orbital wave function Ψ for an electron in an atom has no physical meaning. It is simply a mathematical function of the coordinates of the electron.

64. Plots of the orbital wave function $\Psi(r)$ and probability density $\Psi^2(r)$ Vs distance r of the electron from the nucleus for 2s orbital:



- For 2s orbital the probability density is maximum at the nucleus and it decreases sharply as we move away from it (which is not possible). Hence plot of probability density $\Psi^2(r)$ Vs distance r of the electron from the nucleus was drawn as shown below.



- For 2s orbital, the probability density first decreases sharply to zero and again starts increasing. After reaching small maxima it decreases again and approaches zero as the value of r increases further.

65. The region where this probability density function reduces to zero is called nodal surfaces or simply nodes.

66. Charge cloud diagrams: In these diagrams, dots represent the electron probability density. The density of the dots in a region represents electron probability density in that region.

67. Boundary surface diagram: In this representation, a boundary surface or contour surface is drawn in space for an orbital on which the value of probability density $\Psi^2(r)$ is constant. However, for a given orbital, only that boundary surface diagram of constant probability density is taken to be good representation of the shape of the orbital which encloses a region or volume in which the probability of finding the electron is very high, say, 90%.

68. Radial nodes: Radial nodes occur when the probability density wave function for the electron is zero on a spherical surface of a particular radius. Number of radial nodes = $n - l - 1$

69. Angular nodes: Angular nodes occur when the probability density wave function for the electron is zero along the directions specified by a particular angle. Number of angular nodes = l

70. Total number of nodes = $n - 1$

71. Degenerate orbitals: Orbitals having the same energy are called degenerate orbitals.

72. The stability of an electron in a multi electron system is because of:

- The repulsive interaction of the electrons in the outer shell with the electrons in the inner shell.
- The attractive interactions of electron with the nucleus. These attractive interactions increase with increase of positive charge (Z_e) on the nucleus.

a. The stability of an electron in multi-electron atom is because total attractive interactions are more than the repulsive interactions.

73. Shielding effect or screening effect: Due to the presence of electrons in the inner shells, the electron in the outer shell will not experience the full positive charge on the nucleus.

So due to the screening effect, the net positive charge experienced by the electron from the nucleus is lowered and is known as effective nuclear charge.

Effective nuclear charge experienced by the orbital decreases with increase of azimuthal quantum number (l).

74. Orbitals have different energies because of mutual repulsion between electrons in a multi- electron atom.
75. Orbitals with lower value of $(n+l)$ are filled first as they have lower energy.
76. If two orbitals have the same value of $(n+l)$ then orbital with lower value of n will have lower energy.
77. Energies of the orbitals in the same subshell decrease with increase in atomic number.
78. Filling of electrons: The filling of electrons into the orbitals of different atoms takes place according to Aufbau principle, Pauli's exclusion principle, the Hund's rule of maximum multiplicity
79. Aufbau Principle: In the ground state of the atoms, the orbitals are filled in order of their increasing energies. The order in which the orbitals are filled is as follows:
 $1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 4f, 5d, 6p, 7s...$
It is based on $(n+l)$ rule. It states that the orbital with lower value of $(n+l)$ has lower energy.
80. Pauli Exclusion Principle: No two electrons in an atom can have the same set of four quantum numbers. Only two electrons may exist in the same orbital and these electrons must have opposite spin.
81. Hund's rule of maximum multiplicity: Pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.
82. Electronic configuration of atoms: The electronic configuration of different atoms can be represented in two ways.
 - a. **$s^a p^b d^c$ notation:** In the first notation, the subshell is represented by the respective letter symbol and the number of electrons present in the subshell is depicted, as the super script, like a, b, c, \dots etc. The similar subshell represented for different shells is differentiated by writing the principal quantum number before the respective subshell.
 - b. **Orbital diagram:** In the second notation, each orbital of the subshell is represented by a box and the electron is represented by an arrow (\uparrow) a positive spin or an arrow (\downarrow) a negative spin.

83. Stability of completely filled and half filled subshells:
 - a. Symmetrical distribution of electrons
 - b. Exchange energy

Class: XI
Chapter 3: Classification of Elements
Chapter Notes

Top Concepts

1. Johann Dobereiner classified elements in group of three elements called triads.
2. In Dobereiner's triad the atomic weight of the middle element is very close to the arithmetic mean of the other two elements.
3. Dobereiner's relationship is referred as Law of triads.
4. Since Dobereiner's Law of triads worked only for few elements, it was dismissed.
5. Chancourtois arranged elements in order of increasing atomic weights and made a cylindrical table of elements.
6. John Newland arranged the elements in the increasing order of atomic weight and noted that the properties of the every eighth element are similar to the first one. This relationship is called as "Law of octaves"
7. Lothar Meyer proposed that on arranging the elements in order of increasing atomic weights similarities appear at a regular interval in physical and chemical properties.
8. According to Mendeleev's periodic law the physical and chemical properties of elements are periodic functions of their atomic weights.
9. Merits of Mendeleev's periodic table:
 - Mendeleev's periodic table was very helpful in remembering and studying the properties of large number of elements
 - Mendeleev's periodic table helped in correcting the atomic masses of some of the elements like gold, beryllium and platinum based on their positions in the periodic table
 - Mendeleev could predict the properties of some undiscovered elements like scandium, gallium and germanium. By this intuition, he had left gaps for the undiscovered elements while arranging elements in his periodic table.

10. Demerits of Merits of Mendeleev's periodic table:

- Position of hydrogen is not correctly defined in periodic table. It is placed in group I though it resembles both group 1 and 17.
- In certain pairs of elements increasing order of atomic masses was not obeyed. For example argon (Ar, atomic mass 39.9) is placed before potassium (K, atomic mass 39.1)
- Isotopes were not given separate places in the periodic table although Mendeleev's classification is based on the atomic masses.
- Some similar elements are separated and dissimilar elements are grouped together. For example copper and mercury resembled in their properties but had been placed in different groups. On the other hand lithium and copper were placed together although their properties are quite different.
- Mendeleev did not explain the cause of periodicity among the elements.
- Lanthanoids and actinoids were not given a separated position in the table

11. Moseley performed experiments and studied the frequencies of the X-rays emitted from the elements. With these experiments he concluded that atomic number is more fundamental property of an element than its atomic mass.

12. After Moseley's experimental results Mendeleev's periodic law was modified to modern periodic law.

13. According to Modern periodic law the physical and chemical properties of the elements are periodic functions of their atomic numbers.

14. Modern periodic table is also referred to as long form of periodic table

15. Horizontal rows in the periodic table are called periods.

16. Vertical columns in the periodic table are called groups.

17. In the modern periodic table there are 7 periods and 18 groups.

18. The period number corresponds to highest principal quantum number of elements.
19. First period contains 2 elements
20. Second and third period contains 8 elements
21. Fourth and fifth period contains 18 elements
22. Sixth period contains 32 elements
23. In the modern periodic table, 14 elements of both sixth and seventh periods i.e. lanthanoids and actinoids respectively are placed separately at the bottom of the periodic table.
24. Elements with atomic number greater than 92 are called transuranic elements.
25. According to IUPAC, until a new element's discovery is proved and its name is officially recognized it is given a temporary name. This nomenclature is based Latin words for their numbers.
26. The interim names of the newly discovered elements are derived by combining together the roots in order of digits which make up the atomic number and ium is added at the end.
27. Notation for the IUPAC nomenclature of elements

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

28. The distribution of electrons into orbitals of an atom is called its electronic configuration.
29. The electrons in an orbital are filled according to $n+l$ rule.

30. The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.
31. On moving down a group in a periodic table the number of shell increases from 1 to 7
32. Elements in the same group have same number of valence electrons
33. Value of the principal quantum number for the valence or outermost shell gives the period.
34. The first period has principal quantum number $n=1$, contains two elements and corresponds to K-shell.
35. Since K-shell contains only one orbital (1s) it can accommodate two electrons. Thus there are two elements in K-shell.
36. The second period has principal quantum number $n=2$, contains eight elements and corresponds to L-shell.
37. The 4 orbitals filled in second period are one 2s (with 2 electrons) and three 2p (6 electrons).
38. The third period has principal quantum number $n=3$, contains eight elements and corresponds to M-shell.
39. The four orbitals filled in third period are one 3s (2 electrons) and three 3p (6 electrons).
40. The fourth period has principal quantum number $n=4$, contains eighteen elements.
41. The 9 orbitals filled in fourth period are one 4s (2 electrons), five 3d (with 10 electrons) and three 4p (with 6 electrons).
42. Elements from Scandium ($Z=21$) to Zinc ($Z=30$) are called 3d transition series of elements or first transition series.
43. The fifth period has principal quantum number $n=5$, contains eighteen elements.
44. The nine orbitals filled in fifth period are one 5s (2 electrons), five 4d (10 electrons) and three 5p (6 electrons).
45. Elements from Yttrium ($Z=39$) to Cadmium ($Z=48$) are called 4d transition series of elements or second transition series.

46. The sixth period has principal quantum number $n=6$, contains 32 elements.
47. The 16 orbitals filled in sixth period are one 6s (2 electrons), seven 4f (14 electrons), five 5d (10 electrons) and three 6p (6 electrons).
48. The orbitals filled in seventh period are 7s,5f,6d and 7p
49. Elements from lanthanum ($Z=57$), Hafnium ($Z=72$) to mercury ($Z=80$) are called 5d transition series of elements or third transition series.
50. Fourteen elements from Cerium ($Z=58$) to Lutetium ($Z=71$) are called elements of inner transition series or lanthanoid series.
51. Fourteen elements from Thorium ($Z=90$) to Lawrencium ($Z=103$) are called elements of 5f inner transition series or actinoid series.
52. The 4f and 5f series of elements are placed separately in periodic table to provide a theoretical justification for periodicity occurring at regular intervals.
53. The modern periodic table is divided into four main blocks – s -block, p-block, d-block and f-block depending on the type of orbital that are being filled with exception of hydrogen and helium.
54. The elements in which last electron enter the s-orbital of their outermost energy level are called s-block elements.
55. The s-block consists of two groups, Group-1 and Group-2.
56. The elements of Group-1 are called alkali metals and have ns^1 as the general outer electronic configuration.
57. The elements of Group-2 are called alkaline earth metals and have ns^2 as the general outer electronic configuration.
58. The elements in which last electron enter the p-orbital of their outermost energy level are called p-block elements.
59. The p-block elements constitute elements belonging to group 13 to 18.
60. Elements of s-block and p-block are collectively called representative element
61. The outermost electronic configuration of p-block elements varies from ns^2np^1 to ns^2np^6

62. Elements of group 18 having ns^2np^6 configuration are called noble gases.
63. Elements of group 17 are called halogens
64. Elements of group 16 are called chalcogens
65. Number of valence electrons in group = Group number -10 for elements belonging to group 13 to 18
66. Elements in which the last electron enters d-orbitals of penultimate energy level constitute d-block elements.
67. Elements of group 3 to 12 in the centre of periodic table constitute the d-block elements
68. General outer electronic configuration of d-block elements is $(n-1)d^{1-10} ns^{1-2}$
69. d-block elements constitute transition series elements. The name "transition series" is derived from the fact the d-block elements represent transition in character from reactive metals (belonging to group 1 and 2 constituting s-block) on one side of the periodic table to non-metals (belonging to group 13 to 18 constituting p-block) on other side of the periodic table .
70. Elements in which last electron enters f-orbitals are called f-block elements
71. Elements of Lanthanoid series have general outer electronic configuration of $4f^{1-14} 5d^{0-1} 6s^2$
72. Elements of Actinoid series have general outer electronic configuration of $5f^{1-14} 6d^{0-1} 7s^2$
73. Elements in lanthanoid and actinoid series are called inner transition series.
74. Metals comprise more than 78 % of all known elements and appear on left hand side of periodic table
75. Non-metals are placed on right hand side of periodic table
76. Metals are characterized by having a tendency to lose electron
77. Non-metals are characterised by having tendency to gain electron
78. In general metallic character increases down the group and decreases along period

79. In general non-metallic increases along a period and increases along group
80. Elements showing properties of both metals and non-metals are called metalloids or semi-metals
81. The recurrence of similar properties of elements after certain regular intervals when they are arranged in order of increasing atomic number is called periodicity.
82. The cause of periodicity of properties of elements is due to the repetition of similar electronic configuration of their atoms in the outermost energy shell after certain regular interval.
83. Covalent radius for a homonuclear molecule is defined as one half of the distance between the centres of nuclei of two similar atoms bonded by single covalent bond.
84. For heteronuclear molecule covalent radius may be defined as the distance between the centre of nucleus of atom and mean position of the shared pair of electrons between the bonded atoms.
85. Metallic radius is defined as the one half of the internuclear distance two neighbouring atoms of a metal in a metallic lattice.
86. For simplicity term atomic radius is used for both covalent and metallic radius depending on whether element is non-metal or a metal.
87. Atomic radius decrease with increase in atomic number on going from left to right in a period.
88. Atomic radius of elements increase from top to bottom in a group.
89. van der Waals radius is half of the distance between two similar atoms in separate molecules in a solid.
90. Ionic radius may be defined as the effective distance from the nucleus of the ion upto which it has an influence in the ionic bond.
91. A cation is smaller than the parent atom.
92. An anion is larger than the parent atom.
93. On moving from top to bottom in a group in a periodic table ionic radius increases.

94. On moving from left to right in a period in a periodic table ionic radius decrease.
95. Atoms or ions which contain same number of electrons are called isoelectronic species.
96. In case of isoelectronic cations, the cation with a greater positive charge will have a smaller radius because of greater attraction of electrons to nucleus.
97. In case of isoelectronic anions, the anions with a greater negative charge will have a larger radius because repulsion of electrons will outweigh the nuclear charge.
98. A quantitative measure of the tendency of an element to lose electron is given by ionization enthalpy. It represents the energy required to remove an electron from an isolated gaseous atom in ground state.
99. The energy required to remove second most loosely bound electron is called second ionization energy. The value of second ionization enthalpy is higher than first ionization enthalpy because it is more difficult to remove an electron from a positively charged ion than a neutral atom.
100. The effective nuclear charge experienced by the valence electron in an atom will be less than actual nuclear charge on nucleus because of shielding or screening of valence electron from the nucleus by inner core electrons.
101. On moving from left to right along a period in periodic table ionisation enthalpy increases. On moving along a period successive electrons are added to orbitals in same quantum level and shielding of nuclear charge by inner core of electrons does not increase to an extent to compensate for increased attraction of electron to nucleus. Thus increasing nuclear charge outweighs shielding across a period. Eventually more energy is required to remove outermost electron.
102. On moving from top to bottom in a group in periodic table ionisation enthalpy decreases. On moving down a group successive shells are added and outermost electron move further away from nucleus. Due to electrons present in inner shells shielding of nuclear charge increases. Thus along a group shielding outweighs increasing nuclear charge. Eventually less energy is required to remove an outermost electron.

103. Among various groups in a periodic table, Group 18 elements have highest ionization enthalpy and because of stable electronic configuration
104. When an electron is added to neutral gaseous atom to convert it into a negative ion, the enthalpy change accompanying the process is called electron gain enthalpy.
105. Group 17 elements have high negative electron gain enthalpy because they can attain a stable electronic configuration as of noble gases by accepting an electron.
106. In general electron gain enthalpy becomes more negative from left to right in a period. The effective nuclear charge increases from left to right across a period. Thus it is easier to add to add an electron to a smaller atom because the added electron would be on an average closer to positively charged nucleus.
107. In general electron gain enthalpy becomes less negative as we go from top to bottom in a group. This is because added electron would be farther away nucleus.
108. Elements like O and F have less negative electron gain enthalpy than the succeeding elements like S and Cl respectively of the same group. This is because in case of O and F electron is added to smaller quantum number ($n=2$) and suffers greater repulsion from electrons present in that level. In case of succeeding elements like S and Cl electron is added to $n=3$. The added electron occupies large region of space and experiences less repulsion from electrons in that level.
109. Ability of an atom in a chemical compound to attract shared electrons to itself is called electronegativity.
110. The electrons present in the outermost shell are called valence electrons and these electrons determine the valence of atom.
111. Valence of representative element is usually equal to
- The number of electrons in the valence shell
 - $8 - (\text{the number of electrons in the valence shell})$
112. On moving down a group since the number of valence electrons remains the same, all elements exhibit same valence.

113. Oxidation state of an element in a particular compound gives the charge acquired by its atoms on basis of electronegativity consideration from other atoms in the molecule.
114. It is observed that some elements of second period show similarities with elements of third period present diagonally to each other though belonging to different group. This similarity in properties of elements present diagonally is called diagonal relationship.
115. Lithium is diagonally related to Magnesium, beryllium is diagonally related to aluminium and boron is diagonally related to silicon.
116. The anomalous behavior of first element of s and p block elements of each group as compared to other group members is due to following reasons:
- Small size of atom
 - Large charge/radius ratio
 - High electronegativity
 - Non availability of d-orbitals in their valence shell
117. In second period, the first element of each group has 4 valence orbital (2s and 2p) which are available for bonding. Therefore covalence of first member of each group is only 4.
118. Elements of p-block in the second period displays greater ability to form $p\pi - p\pi$ multiple bond to itself (e.g., $C=C$, $C\equiv C$, $N=N$, $N\equiv N$) and to other second period elements (e.g. $C=O$, $C=N$, $C\equiv N$, $N=O$) compared to subsequent members of the same group.
119. In a periodic table there is high chemical reactivity at two extreme ends and lowest in centre. Maximum chemical reactivity at extreme left (alkali metals) is exhibited by the easy loss of electrons forming a cation and at extreme right (among halogens) shown by gain of electrons forming anion.
120. The elements which readily loose electrons act as strong reducing agent.
121. The elements which readily accept electrons acts as strong oxidizing agent.
122. Tendency of an element to lose or gain electrons is also related to metallic or non-metallic character.

123. Elements in extreme left of the periodic table have a tendency to lose electron and become positively charged. Hence they show metallic character.
124. Elements in extreme right of the periodic table have a tendency to gain electron. Hence they show non-metallic character.
125. Metallic character decreases along a period on moving from left to right in a periodic table.
126. Non-metallic character increases along a period on moving from left to right in a periodic table.
127. Since elements at extreme left of periodic table show metallic character, oxides formed by them are basic.
128. Since elements at extreme right of periodic table show non-metallic character, oxides formed by them are acidic.
129. Oxides of elements in centre of periodic table are amphoteric or neutral.
130. Transition metals of 3d series are less electropositive than group 1 and 2 metals. This is because their size is small as compared to group 1 and 2 elements accompanied by higher ionization enthalpy as compared to group 1 and 2 elements.

Points to Remember

Class: XI

Chapter Name: Chemical Bonding and Molecular Structure

Top Concepts

1. The attractive force which holds together the constituent particles (atoms, ions or molecules) in chemical species is known as chemical bond.
2. Tendency or urge atoms of various elements to attain stable configuration of eight electrons in their valence shell is cause of chemical combination.
3. The principle of attaining a maximum of eight electrons in the valence shell or outermost shell of atoms is known as octet rule.
4. Electronic Theory: Kossel-Lewis approach to chemical Bonding: Atoms achieve stable octet when they are linked by chemical bonds. The atoms do so either by transfer or sharing of valence electrons. Inner shell electrons are not involved in combination process.
5. Lewis Symbols or electron dot symbols: The symbol of the element represents the whole of the atom except the valence electrons (i.e. nucleus and the electrons in the linear energy shells). The valence electrons are represented by placing dots (.) or crosses (x) around the symbol.
6. Significance of Lewis Symbols: The Lewis symbols indicate the number of electrons in the outermost or valence shell which helps to calculate common or group valence.
7. The common valence of an element is either equal to number of dots or valence electrons in the Lewis symbol or it is equal to 8 minus the number of dots or valence electrons.
8. The bond formed by mutual sharing of electrons between the combining atoms of the same or different elements is called a covalent bond.
9. If two atoms share one electron pair, bond is known as single covalent bond and is represented by one dash (-).

10. If two atoms share two electron pairs, bond is known as double covalent bond and is represented by two dashes (=)

11. If two atoms share three electron pairs, bond is known as triple covalent bond and is represented by three dashes (\equiv).

12. The formal charge of an atom in a polyatomic ion or molecule is defined as the difference between the number of valence electrons in an isolated (or free) atom and the number of electrons assigned to that atom in a Lewis structure. It may be expressed as:

Formal charge
on an atom =
in free atom

$$\left(\begin{array}{l} \text{Number of} \\ \text{valence electrons} \\ \text{in free atom} \end{array} \right) - \left(\begin{array}{l} \text{Number of} \\ \text{nonbonding} \\ \text{(lone pair) electrons} \end{array} \right) - \frac{1}{2} \left(\begin{array}{l} \text{Number of} \\ \text{bonding} \\ \text{(shared) electrons} \end{array} \right)$$

13. Significance of Formal charge: The formal charges help in selection of lowest energy structure from a number of possible Lewis structures for a given molecule or ion. Lowest energy structure is the one which has lowest formal charges on the atoms.

14. Expanded octet: Compounds in which central atom has more than eight electrons around it, atom is said to possess an expanded octet.

15. Exceptions to the Octet Rule:

Hydrogen molecule: Hydrogen has one electron in its first energy shell ($n = 1$). It needs only one more electron to fill this shell, because the first shell cannot have more than two electrons. This configuration ($1s^2$) is similar to that of noble gas helium and is stable. In this case, therefore, octet is not needed to achieve a stable configuration.

Incomplete octet of the central atom: The octet rule cannot explain the formation of certain molecules of lithium, beryllium, boron, aluminum, etc.

(LiCl, BeH₂, BeCl₂, BH₃, BF₃) in which the central atom has less than eight electrons in the valence shell as shown below:

Expanded octet of the central atom: There are many stable molecules which have more than eight electrons in their valence shells. For example, PF₅, has ten; SF₆ has twelve and IF₇ has fourteen electrons around the central atoms, P, S, and I respectively.

Odd electron molecules: There are certain molecules which have odd number of electrons, like nitric oxide, NO and Nitrogen dioxide, NO₂. In these cases, octet rule is not satisfied for all the atoms.

It may be noted that the octet rule is based upon the chemical inertness of noble gases. However, it has been found that some noble gases (especially xenon and krypton) also combine with oxygen and fluorine to form a large number of compounds such as XeF₂, KrF₂, XeOF₂, XeOF₄, XeF₆, etc.

This theory does not account for the shape of the molecules.

It cannot explain the relative stability of the molecule in terms of the energy.

16. General Properties of Covalent Compounds

1. The covalent compounds do not exist as ions but they exist as molecules.
2. The melting and boiling points of covalent compounds are generally low.
3. Covalent compounds are generally insoluble or less soluble in water and other polar solvents. However, these are soluble in non-polar solvents.
4. Since covalent compounds do not give ions in solution, these are poor conductors of electricity in the fused or dissolved state.

5. Molecular reactions are quite slow because energy is required to break covalent bonds.
6. Since the covalent bond is localized in between the nuclei of atoms, it is directional in nature.

17. Co- Ordinate Covalent Bond:

Covalent type bond in which both the electrons in the shared pair come from one atom is called a coordinate covalent bond.

Co- Ordinate Covalent Bond is usually represented by an arrow (\rightarrow) pointing from donor to the acceptor atom.

Co- Ordinate Covalent bond is also called as dative bond, donor – acceptor bond, semi- polar bond or co-ionic bond.

18. The electrostatic force of attraction which holds the oppositely charged ions together is known as ionic bond or electrovalent bond.

19. Ionic compounds will be formed more easily between the elements with comparatively low ionization enthalpy and elements with comparatively high negative value of electron gain enthalpy.

20. A quantitative measure of the stability of an ionic compound is provided by its lattice enthalpy and not simply by achieving octet of electrons around the ionic species in the gaseous state

21. Lattice enthalpy may also be defined as the energy required to completely separating one mole of a solid ionic compound into gaseous ionic constituents.

22. Factor affecting lattice enthalpy:

Size of the ions: Smaller the size of the ions, lesser is the internuclear distance and higher will be lattice enthalpy.

Larger the magnitude of charge on the ions, greater will be the attractive forces between the ions. Consequently, the lattice enthalpy will be high.

23. General Properties of Ionic Compounds

Ionic compounds usually exist in the form of crystalline solids.

Ionic compounds have high melting and boiling points.

Ionic compounds are generally soluble in water and other polar solvents having high dielectric constants.

Ionic compounds are good conductors of electricity in the solutions or in their molten states.

The chemical reactions of ionic compounds are characteristic of the constituent ions and are known as ionic reactions.

In ionic – compounds, each ion is surrounded by oppositely charged ions uniformly distributed all around the ion and therefore, electrical field is non-directional.

24. Bond length: It is defined as the average distance between the nuclei of the nuclei of two bonded atoms in a molecule.

25. Covalent radius is half of the distance between two similar atoms joined by single covalent bond in same molecule.

26. Van der Waals radius is one half of the distance between two similar adjacent atoms belonging to two nearest neighbouring molecules of the same substance in the solid state. It is always larger than covalent radii.

27. Bond angle: It is defined as the average angle between orbitals containing bonding electron pairs around the central atom in a molecule.

28. Bond enthalpy: It is defined as amount of energy required to break one mole of bonds of a particular type between atoms in gaseous state.

29. Bond order: The bond order is defined as the number of bonds between two atoms in a molecule.
30. When a single Lewis structure cannot determine a molecule accurately, concept of resonance is used wherein a number of structures with similar energy, positions of nuclei, bonding and non-bonding pairs of electrons are taken as canonical structures of hybrid which describes molecule accurately.
31. Resonance: When a molecule cannot be represented by a single structure but its characteristic properties can be described by two or more than two structures, then the actual structure is said to be a resonance hybrid of these structure.
32. Polarity of Bonds: In reality no bond is completely covalent or completely ionic.
33. Non-polar covalent bond: When a covalent bond is formed between two similar atoms, the shared pair of electrons is equally attracted by the two atoms and is placed exactly in between identical nuclei. Such a bond is called non-polar covalent bond
34. Molecules having two oppositely charged poles are called polar molecules and the bond is said to be polar covalent bond. Greater the difference in the electro-negativity of the atoms forming the bond, greater will be the charge separation and hence greater will be the polarity of the molecule.
35. Dipole moment is defined as the product of the magnitude of the charge and the distance of separation between the charges.
- Dipole moment (μ) = charge (q) x distance of separation (d)
36. Partial Covalent Character in Ionic Bonds: When two oppositely charge ions A^+ and B^- are brought together; the positive ion attracts the outermost electrons of the negative ion. This results in distortion of

electron clouds around the anion towards the cation. This distortion of electron cloud of the negative ion by the positive ion is called polarization.

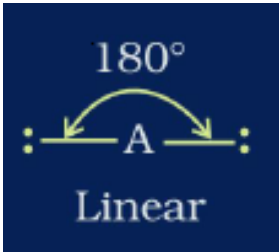
37. Tendency of cation to polarize and polarisability of anion are summarized as Fajan's rules:


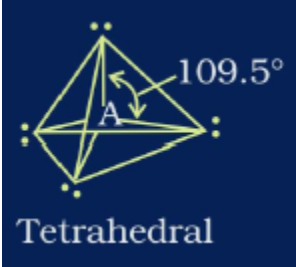


- a.** Smaller the size of the cation, greater is its polarizing power.
- b.** Polarisation increases with increase in size of anion. This is because the electron cloud on the bigger anion will be held less firmly by its nucleus and, therefore, would be more easily deformed towards the cation.
- c.** Larger the charge on cation greater is polarizing power and larger the charge on anion greater is its tendency to get polarized.

38. Valence Shell Electron Pair Repulsion (VSEPR) Theory:

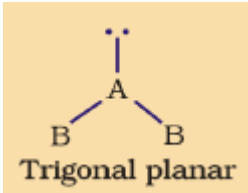
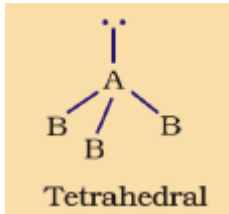
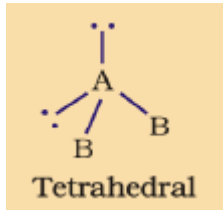
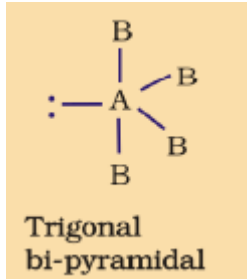
Since Lewis symbols were unable to explain shapes of certain molecules, VSEPR theory was introduced. The basic idea of this theory is that bonded atoms in a molecule adopt that particular arrangement in space around the central atom which keeps them on the average as far apart as possible.

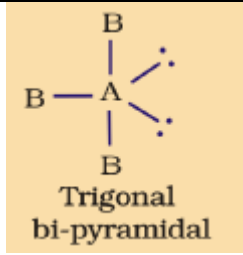
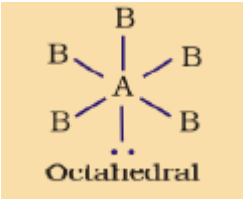
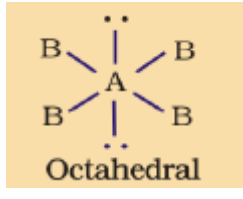
39. Geometry and shapes of molecules in which central atom has no lone pair of electrons

Number of electron pairs	Arrangement of electron pairs	Molecular geometry	Examples
2		Linear	BeCl ₂ , HgCl ₂

3		Trigonal planar	BF ₃
4		Tetrahedral	CH ₄ , NH ₄ ⁺
5		Trigonal bipyramidal	PCl ₅
6		Octahedral	SF ₆

Shapes of simple molecules/Ions with central ions having one or more lone pairs of electrons

Molecule type	No. of bonding pairs	No. of lone pairs	Arrangement of electron pairs	Shape	Example
AB_2E	2	1	 <p>Trigonal planar</p>	Bent	SO_2, O_3
AB_3E	3	1	 <p>Tetrahedral</p>	Trigonal pyramidal	NH_3
AB_2E_2	2	2	 <p>Tetrahedral</p>	Bent	H_2O
AB_4E	4	1	 <p>Trigonal bi-pyramidal</p>	See saw	SF_4

AB_3E_2	3	2	 <p>Trigonal bi-pyramidal</p>	T-shape	ClF_3
AB_5E	5	1	 <p>Octahedral</p>	Square pyramidal	BrF_5
AB_4E_2	4	2	 <p>Octahedral</p>	Square planar	XeF_4

A: Central atom, B is surrounding atoms, E is the lone pair

40. Valence Bond Approach of Covalent Bond

The VSEPR theory gives the geometry of simple molecules but theoretically, it does not explain them and also has limited applications. To overcome these limitations, two important theories based on quantum mechanical principles are commonly used. These are Valence bond (VB) theory and Molecular orbital (MO) theory.

41. Valence Bond Theory

- A discussion of valence bond theory is based on the knowledge of atomic orbitals, electronic configuration of elements, overlap criteria of atomic orbitals and principles of variation and superposition.
- Orbital Overlap Concept of Covalent Bond: When two atoms approach each other, partial merger of two bonding orbitals, known as overlapping of the orbitals occurs.

- Depending upon the type of overlapping, the covalent bonds may be divided as sigma (σ) bond and Pi (π) bond.
- Sigma (σ) bond: This type of covalent bond is formed by the end to end (hand on) overlapping of bonding orbitals along the inter-nuclear axis. The overlap is known as head on overlap or axial overlap. The sigma bond is formed by any one of the following types of combinations of atomic orbitals. Sigma (σ) bond can be formed by $s - s$ overlapping, $s - p$ overlapping, $p - p$ Overlapping etc.
- Pi (π) Bond: This type of covalent bond is formed by the sidewise overlap of the half- filled atomic orbitals of bonding atoms. Such an overlap is known as sidewise or lateral overlap.

42. Hybridization:

- In order to explain characteristic geometrical shapes of polyatomic molecules concept of hybridization is used.
- The process of intermixing of the orbitals of slightly different energies so as to redistribute their energies resulting in the formation of new set of orbitals of equivalent energies and shape.

43. Atomic orbitals used in different types of hybridization.

Shapes of molecules/ions	Hybridisation type	Atomic orbitals	Examples
Linear	sp	one s + one p	BeCl_2
Trigonal planar	sp^2	one s + two p	BCl_3
Tetrahedral	sp^3	one s + three p	CH_4, NH_3
Square planar	dsp^2	one d + one s + two p	$[\text{Ni}(\text{CN})_4]^{2-}$ $[\text{Pt}(\text{Cl})_4]^{2-}$
Trigonal bipyramidal	sp^3d	one s + three p + one d	$\text{PF}_5, \text{PCl}_5$

bipyramidal			
Square pyramidal	sp^3d^2	one s +three p +two d	BrF_5
Octahedral	sp^3d^2 d^2sp^3	one s +three p +two d two d + one s +three p	$SF_6, [CrF_6]^{3-}, [Co(NH_3)_6]^{3+}$

44. Molecular Orbital Theory (MOT):

- Basic idea of MOT is that atomic orbitals of individual atoms combine to form molecular orbitals. Electrons in molecule are present in the molecular orbitals which are associated with several nuclei.
- The molecular orbital formed by the addition of atomic orbitals is called the bonding molecular orbital (σ).
- The molecular orbital formed by the subtraction of atomic orbital is called antibonding molecular orbital (σ^*).
- The sigma (σ) molecular orbitals are symmetrical around the bond-axis while pi (π) molecular orbitals are not symmetrical.
- Sequence of energy levels of molecular orbitals changes for diatomic molecules like $Li_2, Be_2, B_2, C_2, N_2$ is $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$
- Sequence of energy levels of molecular orbitals changes for diatomic molecules like O_2, F_2, Ne_2 is $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$

Bond order (b.o.) is defined as one half the difference between the number of electrons present in the bonding and the antibonding orbitals

45. Hydrogen Bonding:

- The attractive force which binds hydrogen atom of one molecule with electronegative atom like F, O or N of another molecule is known as hydrogen bond or hydrogen bonding.
- Magnitude of hydrogen bonding is maximum in solid state and least in gaseous state.

- Intermolecular hydrogen bond is formed between two different molecules of same or different substances.
- Intramolecular hydrogen bond is formed between the hydrogen atom and highly electronegative like O, F or N present in the same molecule.

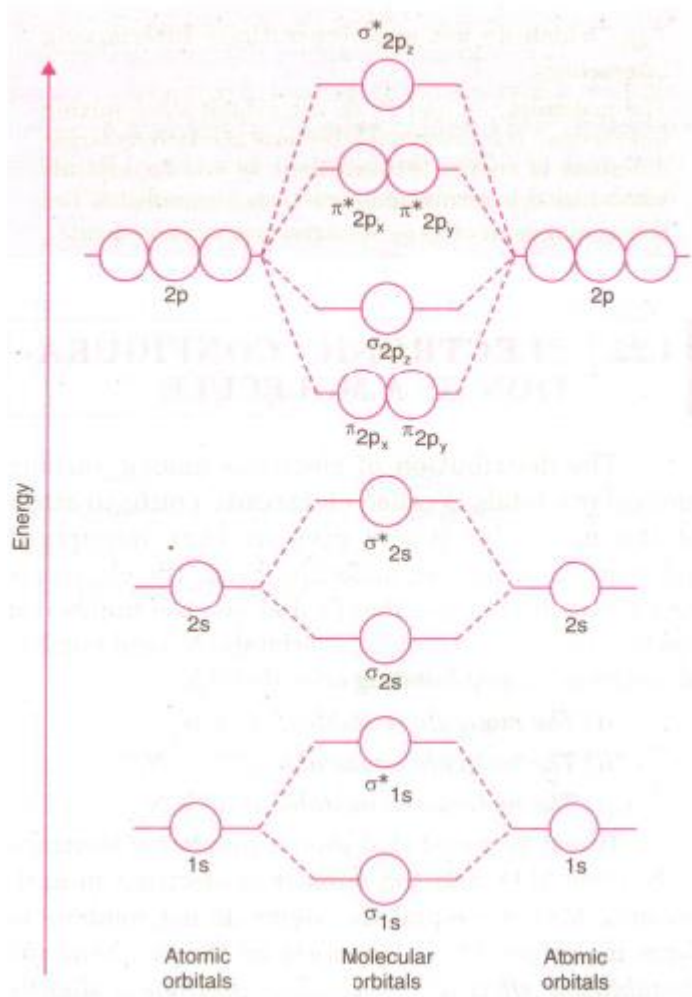
Top Formula:

1.

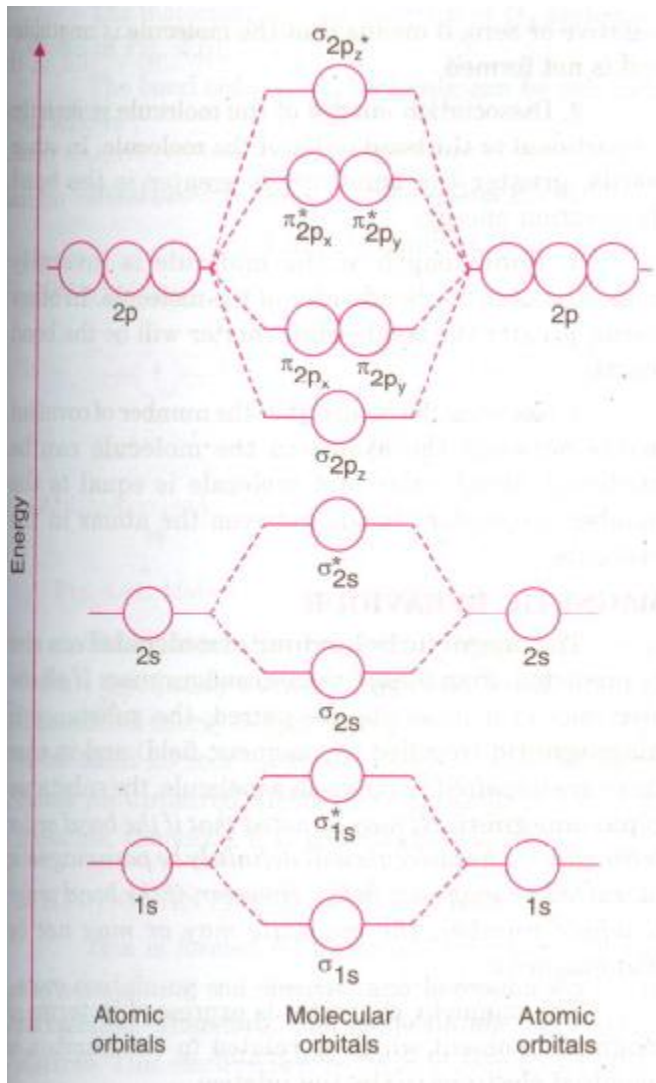
$$\% \text{ ionic character} = \frac{\text{Observed dipole moment}}{\text{Dipole moment for complete ionic character}} \times 100$$

Top Diagrams

Molecular orbital energy level diagram for Li_2 , Be_2 , B_2 , C_2 , N_2 molecules



Molecular orbital energy level diagram for O_2 , F_2 , Ne_2 etc.



Points to Remember**Subject: Chemistry****Class: XI****Chapter: States of matter****Top concepts**

1. Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms and molecules).
2. Attractive intermolecular forces are known as van der Waals forces which include dispersion forces or London forces, dipole – dipole forces, and dipole – induced dipole forces.
3. Dispersion forces or London forces exist between neutral atoms like that of noble gases or non – polar molecules like O₂, N₂, H₂ etc. Magnitude of dispersion forces depends on the polarisability of the neutral atom or molecule.
4. Dipole – Dipole forces occur between the molecules have permanent dipole such as HCl, NH₃, H₂O etc.
5. Dipole – Induced Dipole forces operate between the polar molecule having permanent dipole and the molecule having no permanent dipole.
6. Hydrogen bond is a special case of dipole-dipole interaction. When hydrogen atom is bonded to atoms of highly electronegative elements such as fluorine, oxygen or nitrogen, hydrogen atom forms a weak bond with the electronegative atom of the other molecule. This weak bond is called hydrogen bond.
7. Thermal energy is the energy of a body arising from motion of its atoms or molecules.
8. Three states of matter are the result of balance between intermolecular forces and the thermal energy of the molecules.

9. Predominance of molecular interactions result into change of gases to liquid to solid state while predominance of thermal energy results into change of solid to liquid to gas.

10. Physical properties of gas

- Gases are highly compressible
- Gases exert pressure equally in all directions
- Gases have density much lower than the solids and liquids
- Gases do not have definite volume and the shape. They assume volume and shape of the container
- Gases mix evenly and completely in all proportions without any mechanical aid.

11. According to Boyle's law at constant temperature, the pressure of a fixed amount (i.e., number of moles n) of gas varies inversely with its volume. Mathematically, it is written as

$$p \propto \frac{1}{V} \quad (n, T \text{ are constant})$$

or $pV = k_1$ (constant)

Value of k_1 depends on amount of gas and temperature and units in which p and V are expressed

12. Boyle's law expresses in a quantitative manner that gases are compressible because when a given mass of a gas is compressed, the same number of molecules occupy a smaller space.

13. According to Charles' Law at constant pressure, the volume of fixed mass of a gas is directly proportional to absolute temperature. Mathematically, it is written as

$$V \propto T \quad (n, P \text{ are constant})$$

$$\text{or } \frac{V}{T} = k_2 \text{ (constant)}$$

Value of k_2 depends on amount of gas and pressure and units in which volume V is expressed

14. The lowest hypothetical or imaginary temperature at which gases are supposed to occupy zero volume is called Absolute zero.

15. According to Gay Lussac's Law at constant volume, pressure of a fixed amount of a gas varies directly with the temperature. Mathematically, it is written as

$$p \propto T \quad (n, V \text{ are constant})$$

$$\text{or } \frac{p}{T} = k_3 \text{ (constant)}$$

16. Standard temperature and pressure means 273.15 K (0°C) temperature and 1 bar (i.e., exactly 10^5 Pascal) pressure.

17. According to Avogadro Law equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules. Mathematically, it is written as

$$V \propto n \quad (p, T \text{ are constant})$$

$$\text{or } \frac{V}{n} = k_4 \text{ (constant)}$$

18. A gas that follows Boyle's Law, Charles' law and Avogadro law strictly is called an ideal gas.

19. Ideal Gas Equation : $pV = nRT$

R is called gas constant. It is same for all gases. Therefore it is also called Universal Gas Constant.

20. Different values of Universal Gas Constant

$$R = 8.314 \text{ Pa m}^3 \text{ K}^{-1} \text{ mol}^{-1}$$

$$= 8.314 \times 10^{-2} \text{ bar L K}^{-1} \text{ mol}^{-1}$$

$$= 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$$

21. Combined Gas Law : $\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$

22. According to Dalton's Law of Partial Pressures the total pressure exerted by the mixture of non - reactive gases is equal to the sum of the partial pressures of individual gases. Mathematically, it is written as

$$p_{\text{total}} = p_1 + p_2 + p_3 + \dots \text{ (T, V are constant)}$$

p_{total} is the total pressure exerted by the mixture of gases

p_1, p_2, p_3 etc. are partial pressures of gases

23. Partial pressure in terms of mole fraction is expressed as

$$p_i = x_i \times p_{\text{total}}$$

p_i is partial pressure of i th gas

x_i is mole fraction of i th gas

24. Assumptions or postulates of Kinetic Molecular Theory of Gases:

- Gases consist of large number of minute identical particles (atoms or molecules) which are very small.
- Gas molecules are so far apart from each other that the actual volume of the molecules is negligible as compared to the total volume of gas. They are thus considered as point masses.
- There is no force of attraction between the particles of a gas at ordinary temperature and pressure.
- Particles of a gas are always in constant and random motion.
- Particle of a gas move in all possible directions in straight lines. During their random motion, they collide with each other and with the walls of the container. Pressure is exerted by the gas as a result of collision of the particles with the walls of the container.
- Collisions of gas molecules are perfectly elastic. This means that total energy of molecules before and after the collision remains same.
- At any particular time, different particles in the gas have different speeds and hence different kinetic energies.

- In kinetic theory it is assumed that average kinetic energy of the gas molecules is directly proportional to the absolute temperature.
25. Gases show deviation from ideal behaviour because of two faulty assumptions
- There is no force of attraction between the molecules of a gas.
 - Volume of the molecules of a gas is negligibly small in comparison to the space occupied by the gas.
26. At low temperature and high pressure gases deviate from ideal behaviour i.e. gases behaves as real gases
27. At low pressure and high temperature gases show ideal behaviour i.e. gases behaves as ideal gases
28. van der Waals equation of state is given below
- $$\left(p + \frac{an^2}{V^2}\right)(V - nb) = nRT$$
- a, b are van der Waals constant
n is the number of moles of gas
29. The deviation from ideal behaviour can be measured in terms of compressibility factor Z, which is the ratio of product pV and nRT.
- Mathematically: $Z = \frac{pV}{nRT} = \frac{V_{\text{real}}}{V_{\text{ideal}}}$
- For ideal gas $Z = 1$
For real gas $Z > 1$ or $Z < 1$
30. The temperature at which a real gas obeys ideal gas law over an appreciable range of pressure is called Boyle temperature or Boyle point.
31. Liquefaction of Gases: Gases can be liquefied by lowering temperature and increasing pressure.
32. Critical temperature (T_c) is the temperature above which a gas cannot be liquefied however high pressure may be.

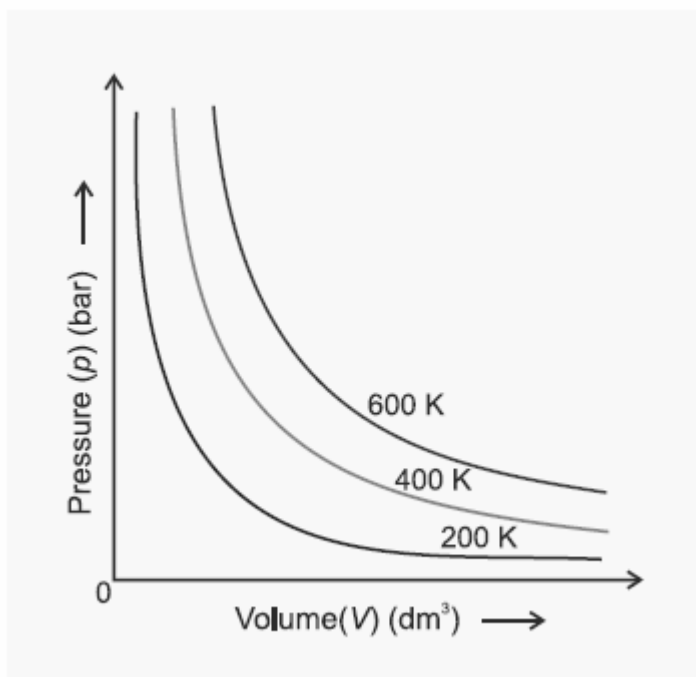
33. Critical pressure (p_c) is minimum pressure required to liquefy gas at the critical temperature.
34. Critical volume (V_c) is the volume occupied by one mole of gas at critical temperature and critical pressure.
35. The temperature at which vapour pressure of liquid is equal to the external pressure is called boiling temperature at that pressure. At 1 atm pressure boiling temperature is called normal boiling point. If pressure is 1 bar then the boiling point is called standard boiling point of the liquid.
36. Standard boiling point of the liquid is slightly lower than the normal boiling point because 1 bar pressure is slightly less than 1 atm pressure.
37. Surface tension (γ) is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid. SI unit is expressed as N m^{-1} .
38. Surface tension decreases with rise in temperature.
39. Coefficient of viscosity (η) is the force when velocity gradient is unity and the area of contact is unit area. SI unit of coefficient of viscosity is N s m^{-2} or $\text{kg m}^{-1} \text{s}^{-1}$. In cgs system the unit of coefficient of viscosity is poise
40. Viscosity of liquids decreases as the temperature rises because at high temperature molecules have high kinetic energy and can overcome the intermolecular forces to slip past one another between the layers.
41. In Boyle's law pressure versus volume graph is plotted at constant temperature. Such a graph is called as an isotherm because temperature remains constant
42. In Charles law volume versus temperature graph is plotted at constant pressure. Such a graph is called as isobar because pressure remains constant.

43. In Gay Lussac's law pressure versus temperature graph is plotted at constant volume. Such a graph is called as an isochore because volume remains constant.

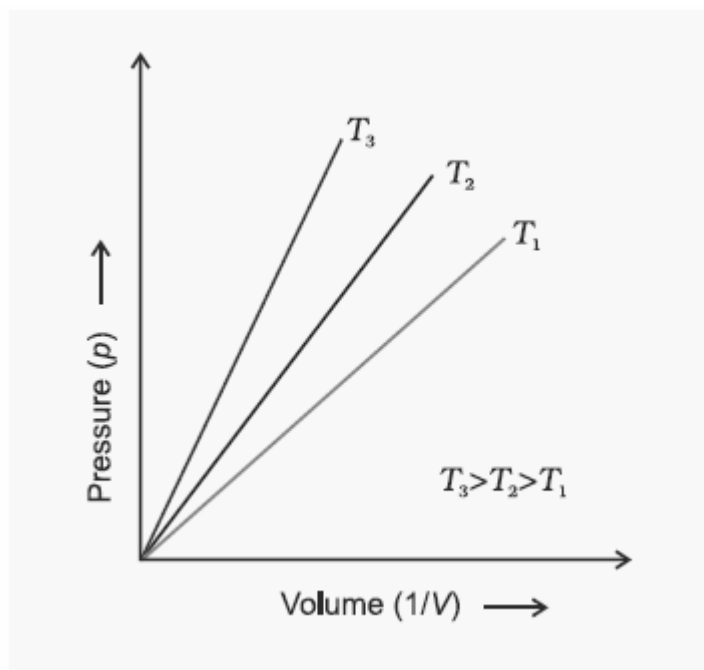
Top Diagrams

1. Representation of Boyle's law graphically:

Graph 1: Plot of pressure (p) versus volume of gas (V) at different temperatures

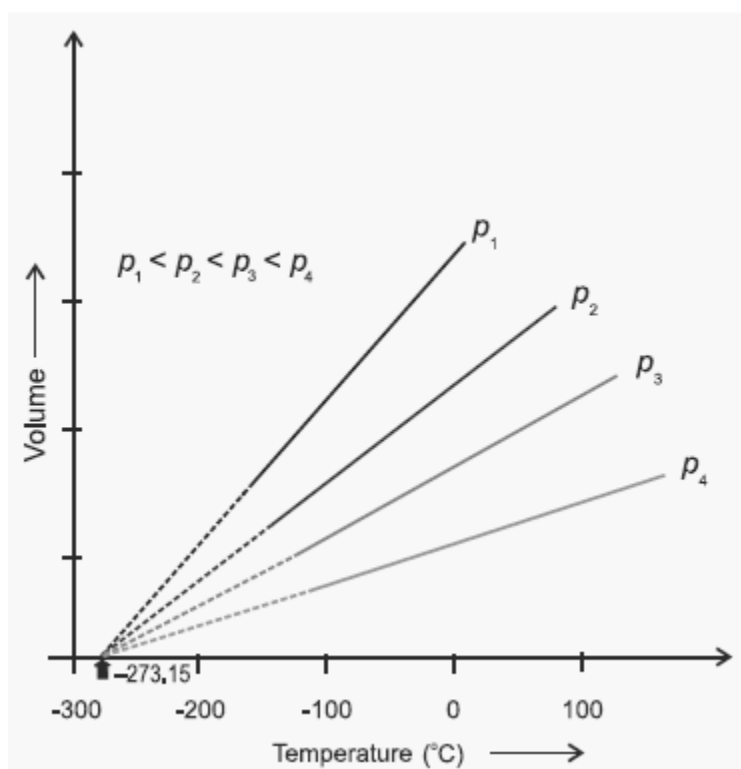


Graph 2: Plot of pressure (p) versus ($1/V$) at different temperatures



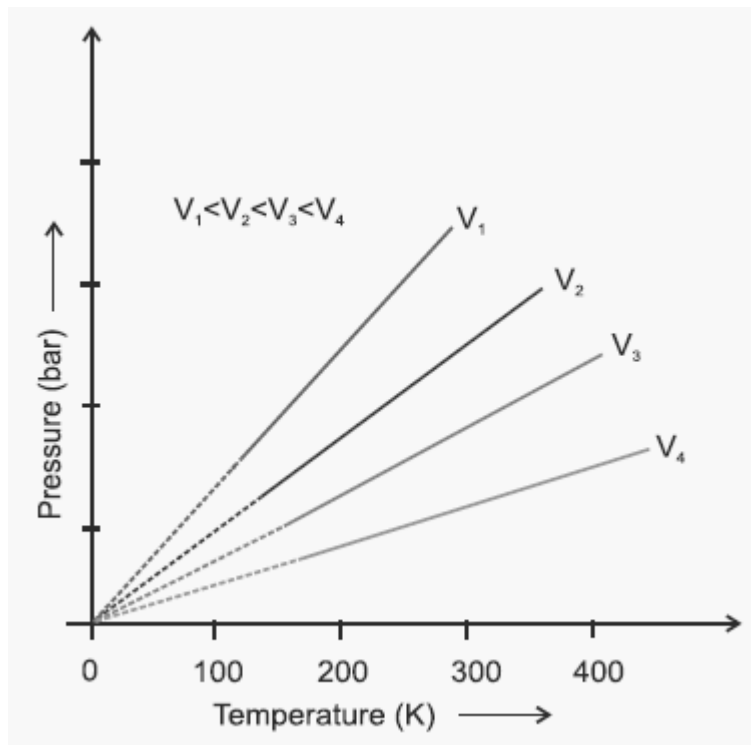
2. Representation of Charles law graphically

Plot of volume versus temperature at different pressures



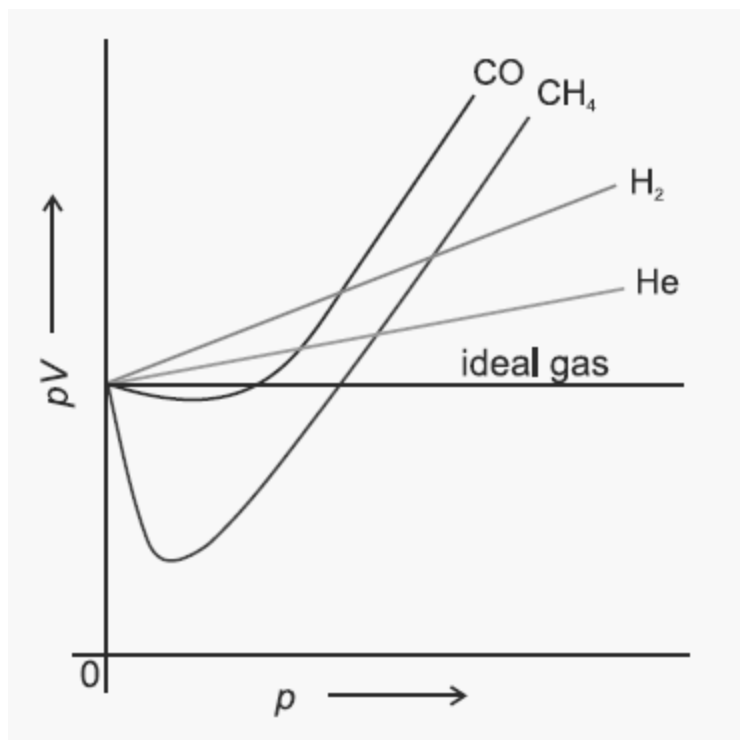
3. Representation of Charles law graphically

Plot of pressure versus temperature at different volume

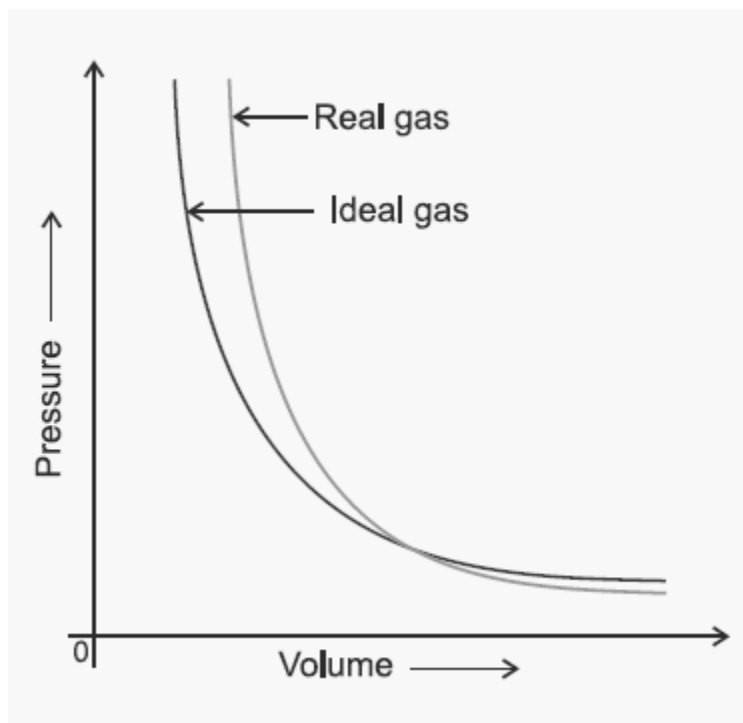


4. Deviation from ideal gas behaviour

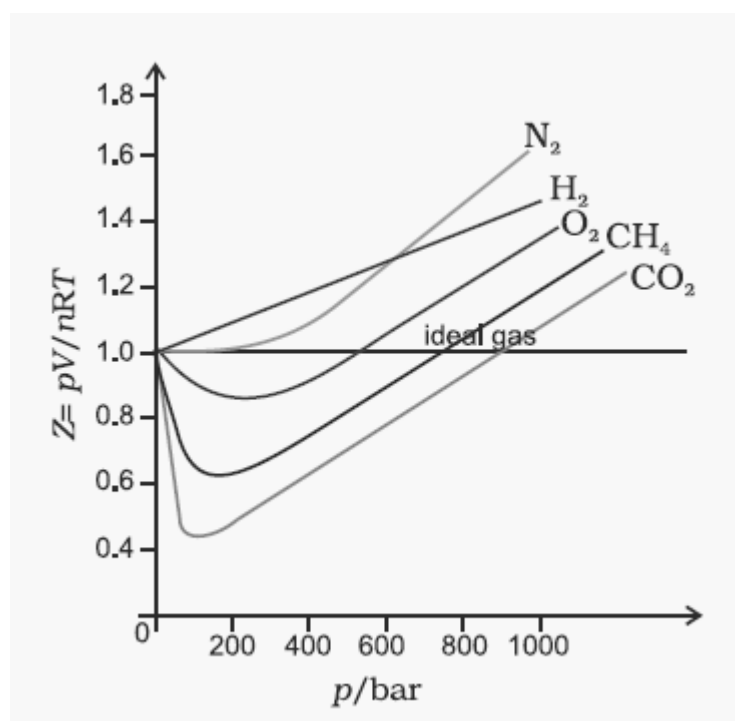
Graph 1: Plot of pressure-volume (pV) versus pressure (p) at constant temperature for ideal and real gases



Graph 2: Plot of pressure (p) versus volume (V) for ideal and real gases



5. Plot of compressibility factor (Z) versus pressure(p) for some gases



Chapter Notes

Subject: Chemistry

Class: XI

Chapter: Thermodynamics

Top concepts

1. The branch of science which deals with study of different forms of energy and their interconversion is called thermodynamics.
2. A system in thermodynamics refers to that part of universe in which observations are made.
3. The remaining portion of universe which is not part of system constitutes the surroundings. The surroundings include everything other than the system.
4. The wall (real or imaginary) that separates the system from the surroundings is called boundary.
5. Types of the System

Types of the System	Exchange of energy	Exchange of matter
Open	Yes	Yes
Closed	Yes	No
Isolated	No	No

6. The state of a thermodynamic system is described by its measurable or macroscopic (bulk) properties. The state of the surroundings can never be completely specified.
7. State variables are the measurable properties of system required to describe the state of the system. Examples are temperature, pressure, volume etc.
8. Various types of processes

Type of process	Definition
Isothermal	Process in which temperature of system remains constant
Adiabatic	Process in which there is no transfer of heat between the system and surroundings
Isobaric	Process in which pressure of system

	remains constant
Isochoric	Process in which volume of system remains constant

9. Every substance possesses definite amount of energy which depends on factors such as chemical nature, temperature and pressure etc. Internal energy, U of the system represents the total energy of the system

10. Work done in a adiabatic process and sign convention

Work done (w_{ad})	Sign of w
By the system	-
On the system	+

11. Heat changes and sign convention

Heat(q)transferred	Sign of q
From surroundings to the system	+
From system to the surroundings	-

12. First Law of Thermodynamics: $\Delta U = q + w$
13. First law of thermodynamics is also called as law of conservation of energy i.e. energy can neither be created nor destroyed. It also states that the energy of an isolated system is constant.
14. Absolute value of the internal energy can never be determined only change in internal energy can be determined.
15. A process or change is said to be reversible, if a change is brought out in such a way that the process could, at any moment, be reversed by an infinitesimal change.
16. For an isothermal reversible process

$$q = -w_{rev} = -2.303nRT \log \frac{V_f}{V_i}$$

17. For isothermal irreversible change $q = -w = p_{ex}(V_f - V_i)$
18. Expansion of a gas in vacuum ($p_{ex} = 0$) is called free expansion. No work is done during free expansion of an ideal gas whether the process is reversible or irreversible
19. For adiabatic change, $q = 0$ and $\Delta U = w_{ad}$

20. The enthalpy of a system may be defined as the sum of the internal energy and the product of its pressure and volume. It is denoted by the symbol H and is given by $H = U + PV$

21. Change in enthalpy $\Delta H = \Delta U + p\Delta V$

22. Enthalpy change and sign convention

Type of reaction	Definition	ΔH
Exothermic reactions	Reactions in which heat is evolved during the reaction	-
Endothermic reactions	Reactions in which heat is absorbed during the reaction	+

23. An extensive property is a property whose value depends on the quantity or size of matter present in the system. Examples are mass, volume, internal energy, enthalpy, heat capacity, etc

24. An intensive property is a property whose value does not depend on the quantity or size of matter present in the system. Examples are temperature, density, pressure etc.

25. Specific heat, also called specific heat capacity is the quantity of heat required to raise the temperature of one unit mass of a substance by one degree Celsius (or one Kelvin).

26. Relation between amount of heat, q , required to raise the temperatures of a sample, specific heat of the substance, c and change in temperatures is given as $q = mc \Delta T = C\Delta T$

27. At constant pressure as $q_p = C_p \Delta T = \Delta H$

28. At constant volume as $q_v = C_v \Delta T = \Delta U$

29. For a mole of an ideal gas, $C_p - C_v = R$

30. Calorimetry is an experimental technique that helps determining energy changes associated with chemical or physical processes

31. The enthalpy change accompanying a reaction is called the reaction enthalpy ($\Delta_r H$)

$\Delta_r H = (\text{sum of enthalpies of products}) - (\text{sum of enthalpies of reactants})$

32. The standard enthalpy of reaction is the enthalpy change for a reaction when all the participating substances are in their standard states. The standard state of a substance at a specified temperature is its pure form at pressure of 1 bar.

33. Different types of enthalpy

No.	Type of enthalpy	Definition	Symbol
1	Standard Enthalpy of fusion or Molar enthalpy of fusion	The enthalpy change that accompanies melting of one mole of a solid substance in standard state	$\Delta_{\text{fus}} H^\theta$
2	Standard Enthalpy of vaporization or Molar enthalpy of vaporization	The enthalpy change that accompanies vaporizing of one mole of a liquid at constant temperature and under standard pressure (1bar)	$\Delta_{\text{vap}} H^\theta$
3	Standard Enthalpy of sublimation	The enthalpy change when one mole of a solid substance sublimates at a constant temperature and under standard pressure (1bar)	$\Delta_{\text{sub}} H^\theta$
4	Enthalpy of formation or Standard molar enthalpy of formation	The enthalpy change for the formation of one mole of a compound from its elements in their most stable states of aggregation	$\Delta_f H^\theta$
5	Enthalpy of combustion	Enthalpy change when 1 mole of substance is completely burnt in excess of oxygen or air	$\Delta_{\text{comb}} H^\theta$ or $\Delta_c H^\theta$
7	Enthalpy of atomisation	Enthalpy change in breaking one mole of bonds of a substance completely into atoms in gaseous state	$\Delta_a H^\theta$
8	Bond dissociation enthalpy	Amount of energy required to break one mole of bond of a particular type between atoms in gaseous state	$\Delta_{\text{bond}} H^\theta$
9	Enthalpy of	Enthalpy of solution of a substance is	$\Delta_{\text{sol}} H^\theta$

	solution	the enthalpy change when one mole of substance dissolves in a specified amount of solvent	
--	----------	---	--

34. A balanced chemical equation together with the value of its $\Delta_r H^\theta$ is called a thermochemical equation

35. Hess's Law of Constant Heat Summation: If a reaction takes place in several steps then its standard reaction enthalpy is the sum of the standard enthalpies of the intermediate reactions into which the overall reaction may be divided at the same temperature

36. Born-Haber Cycle is used to determine lattice enthalpy of ionic compounds since they cannot be determined by experiment directly.

37. The lattice enthalpy of an ionic compound is the enthalpy change which occurs when one mole of an ionic compound dissociates into its ions in gaseous state

$$38. \Delta_{\text{sol}} H^\theta = \Delta_{\text{lattice}} H^\theta + \Delta_{\text{hyd}} H^\theta$$

39. A process that has natural tendency of occurrence in a particular direction and is reversible only by application of some external agency is known as spontaneous process.

40. The processes which are forbidden and are made to take place only by supplying energy continuously from outside the system are called non-spontaneous process.

41. Entropy (S) is the measure of randomness of a system. It is a state function.

42. Entropy increases from solid to gas. Entropy is maximum for gases as they have maximum disorder.

43. According to Second law of thermodynamics entropy of the universe always increases during a spontaneous change

44. For reversible process entropy change is given by

$$\Delta S = \frac{q_{\text{rev}}}{T} \text{ where}$$

q_{rev} is heat absorbed or released during the reaction

T is the temperature of the reaction

45. At constant pressure $q_{\text{rev}} = \Delta H$

$$\therefore \Delta S = \frac{\Delta H}{T}$$

46. For spontaneous process $\Delta S(\text{total}) = \Delta S(\text{sys}) + \Delta S(\text{surr}) > 0$

47. Decrease in enthalpy and increase in entropy of system independently cannot serve the criteria for spontaneity

48. Gibbs function or Gibbs energy is denoted by G.

$$G = H - TS$$

49. The criteria for spontaneous reaction in relation to ΔG at constant pressure and constant temperature.

If $\Delta G < 0$, process is spontaneous

If $\Delta G = 0$, process is in equilibrium

If $\Delta G > 0$, process is non-spontaneous

50. Relation between Gibbs energy change ($\Delta_r G^\ominus$) and equilibrium constant (K) $\Delta_r G^\ominus = -RT \ln K$

51. In case of endothermic reactions enthalpy of reaction ($\Delta_r H^\ominus$) is large and positive. In such a case, value of equilibrium constant (K) will be much smaller than 1 and the reaction is unlikely to form much product

52. In case of exothermic reactions enthalpy of reaction ($\Delta_r H^\ominus$) is negative and $\Delta_r G^\ominus$ is also large and negative. In such cases equilibrium constant (K) is larger than 1 and reactions go to completion

53. Effect of temperature on Spontaneity of Reactions

$\Delta_r H^\ominus$	$\Delta_r S^\ominus$	$\Delta_r G^\ominus$	Description*
-	+	-	Reaction spontaneous at all temperature
-	-	- (at low T)	Reaction spontaneous at low temperature
-	-	+ (at high T)	Reaction nonspontaneous at high temperature
+	+	+ (at low T)	Reaction nonspontaneous at low temperature
+	+	- (at high T)	Reaction spontaneous at high temperature
+	-	+ (at all T)	Reaction nonspontaneous at all temperatures

Subject: Chemistry

Class: XI

Chapter: Equilibrium

Top concepts

1. When the reactants in a closed vessel at a particular temperature react to give products, the concentrations of the reactants keep on decreasing, while those of products keep on increasing for some time after which there is no change in the concentrations of either of the reactants or products. This stage of the system is the dynamic equilibrium. At equilibrium rates of the forward and reverse reactions become equal.
2. General characteristics of equilibrium involving physical processes
 - Measurable properties of system becomes constant at equilibrium
 - Equilibrium can be established only in case of closed system
 - Equilibrium is always dynamic in nature i.e. the process does not stop but changes take place in forward and reverse direction with same rate
 - When equilibrium is attained for a physical process, it is characterised by constant value of one of its parameters at a given temperature. For example for equilibrium $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$, the pressure becomes constant
 - The magnitude of the constant value of the concentration related expression gives an indication of the extent to which the reaction proceeds before acquiring equilibrium.
3. During a chemical reaction when the rates of the forward and reverse reactions become equal, the concentrations of the reactants and the products remain constant. This is the stage of chemical equilibrium. Chemical equilibrium is dynamic in nature.
4. General characteristics of equilibrium involving chemical processes
 - Chemical equilibrium is dynamic in nature
 - Properties of system becomes constant at equilibrium and remains unchanged thereafter
 - Equilibrium can be established only in case of closed system
 - Equilibrium can be approached from either direction
 - A catalyst does not alter the equilibrium point

5. Law of mass action: Rate of a chemical reaction is directly proportional to the product of molar concentration of the reactants at a constant temperature at any given time.
6. Equilibrium Law or Law of Chemical Equilibrium: At a given temperature, the product of concentrations of the reaction products raised to the respective stoichiometric coefficient in the balanced chemical equation divided by the product of concentrations of the reactants raised to their individual stoichiometric coefficients has a constant value
7. Equilibrium reactions in which all reactants and products are in same phase are called homogeneous equilibrium reactions.
8. Equilibrium reactions in which all reactants and products are in different phases are called heterogeneous equilibrium reactions
9. For a general reaction of the type:
 $aA + bB \rightleftharpoons cC + dD$

The equilibrium constant may be represented as

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

A and B are reactants

C and D are products

Exponents a, b, c and d have same values as in balanced chemical equation

If the concentration expressed is in molar concentrations then $K = K_c$

i.e. $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

10. For a gaseous reaction of the type:



The equilibrium constant may be represented as

$$K_p = \frac{[p_C]^c [p_D]^d}{[p_A]^a [p_B]^b}$$

p_A, p_B, p_C, p_D are the partial pressures of A, B, C and D respectively

Exponents a, b, c and d have same values as in balanced chemical equation

11. Relation between K_p and K_c

$$K_p = K_c (RT)^{\Delta n}$$

$\Delta n = (\text{number of moles of gaseous products}) - (\text{number of moles of gaseous reactants})$
in the balanced chemical equation.

12. In heterogeneous equilibrium pure solids or liquids are present but their concentrations or partial pressures do not appear in the expression of the equilibrium constant.

13. Characteristic of equilibrium constant

- Equilibrium constants is applicable only when the concentration of reactants and products have attained their equilibrium state
- The value of equilibrium constant is independent of initial concentrations of the reactants and products
- Equilibrium constant has a definite value for every reaction at a particular temperature
- Equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the reaction in the forward direction
- The value of equilibrium constant tells the extent to which a reaction proceeds in forward or reverse reaction
- Equilibrium constant is independent of presence of catalyst

14. Applications of equilibrium constant

1) Predicting the Extent of a Reaction

- If $K_C > 10^3$, products predominate over reactants, i.e., if K_C is very large, the reaction proceeds nearly to completion
- If $K_C < 10^{-3}$, reactants predominate over products, i.e., if K_C is very small, the reaction proceeds rarely
- If K_C is in the range of 10^{-3} to 10^3 , considerable concentrations of both reactants and products are present

2) Predicting the Direction of the Reaction.

Reaction quotient (Q) is the ratio of product of concentration (or partial pressure) of products to that of reactants at any stage of the reaction

- If $Q_c > K_c$, the reaction will proceed in the direction of reactants (reverse reaction).
- If $Q_c < K_c$, the reaction will proceed in the direction of the products (forward reaction)
- If $Q_c = K_c$, the reaction mixture is already at equilibrium.

3) Calculating Equilibrium Concentrations

15. Relation between Equilibrium constant K, Reaction quotient (Q) and Gibbs energy (G)

$$K = e^{-\frac{\Delta G^\theta}{RT}}$$

- If, $\Delta G^\theta < 0$ then $-\frac{\Delta G^\theta}{RT}$ is positive, and $e^{-\frac{\Delta G^\theta}{RT}} > 1$, making $K > 1$. This implies reaction is spontaneous or the reaction proceeds in the forward direction to form products predominantly.
- If, $\Delta G^\theta > 0$ then $-\frac{\Delta G^\theta}{RT}$ is negative, and $e^{-\frac{\Delta G^\theta}{RT}} < 1$, making $K < 1$. This implies reaction is non-spontaneous or the reaction proceeds in the forward direction to such a small degree that only a very minute quantity of product is formed.

16. According to Le Chatelier's principle change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner so as to reduce or to counteract the effect of the change

1) Effect of Concentration Change

Increase in concentration of any of the reactants $\xrightarrow{\text{Shifts the equilibrium to}}$ Forward direction

Increase in concentration of any of the products $\xrightarrow{\text{Shifts the equilibrium to}}$ Backward direction

2) Effect of Pressure Change

Increase in pressure $\xrightarrow{\text{Shifts the equilibrium in direction of}}$ Lesser number of gaseous moles

Decrease in pressure $\xrightarrow{\text{Shifts the equilibrium in direction of}}$ Larger number of gaseous moles

3) Effect of Addition of an Inert Gas: Addition of an inert gas at constant volume has no effect on equilibrium

4) Effect of Temperature

Increase in temperature $\xrightarrow{\text{Shifts the equilibrium in direction of}}$ Endothermic reaction

Decrease in temperature $\xrightarrow{\text{Shifts the equilibrium in direction of}}$ Exothermic reaction

5) Effect of a Catalyst

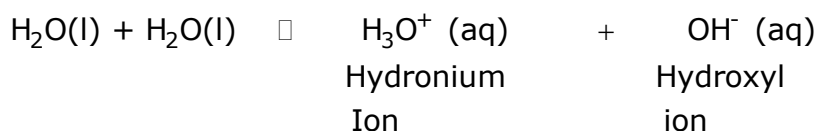
Catalyst does not shift the equilibrium in any direction because it favours the forward and backward reaction equally

17. Substances that conduct electricity in their molten states or in form of their aqueous solutions are called electrolytes

18. Substances which ionize almost completely into aqueous solutions are called strong electrolytes

19. Ionic Equilibrium: Equilibrium established between unionized molecules and ions in solution of weak electrolytes.

20. Substances which ionize to a small extent in aqueous solution are called weak electrolytes
21. Acids, bases and salts which are electrolytes, conduction of electricity by their aqueous solutions is due to anions and cations produced by the dissociation or ionization of electrolytes in aqueous solution.
22. According to Arrhenius theory, acids are substances that dissociates in water to give hydrogen ions $H^+(aq)$ and bases are substances that produce hydroxyl ions $OH^-(aq)$.
23. Limitation of Arrhenius concept of acid and base: The concept is applicable only to aqueous solutions and it does not account for the basicity of substances like, ammonia which do not possess a hydroxyl group.
24. According to Brönsted-Lowry theory, acid is a substance that is capable of donating a hydrogen ion H^+ and bases are substances capable of accepting a hydrogen ion, H^+ .
25. The acid-base pair that differs only by one proton is called a conjugate acid-base pair.
26. Conjugate acid has one extra proton and each conjugate base has one less proton.
27. If Brönsted acid is a strong acid then its conjugate base is a weak base and vice versa.
28. According to Lewis concept an acid as a species which accepts electron pair and base which donates an electron pair.
29. Electron deficient species like $AlCl_3$, Co^{3+} , Mg^{2+} , etc. can act as Lewis acids while species like H_2O , NH_3 , OH^- etc. which can donate a pair of electrons, can act as Lewis bases.
30. Unique ability of water to react with itself to produce a hydronium ion and a hydroxyl ion is called self ionization or autoprotolysis of water.



31. Ionic product of water is the product of concentration of hydronium ions and hydroxyl ions



$$\therefore K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2},$$

K is Equilibrium constant

$$\therefore K[\text{H}_2\text{O}]^2 = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$\therefore K_w = [\text{H}_3\text{O}^+][\text{OH}^-],$$

K_w is ionic product of water

32. In pure water, the concentration of hydronium ion is always equal to the concentration of hydroxyl ion i.e. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

33. The pH of a solution is defined as the negative logarithm to the base 10 of the activity of hydrogen ion i.e. $\text{pH} = -\log [\text{H}^+]$

34. Types of solutions based on pH values

Types of solutions	Hydrogen ion concentration i.e. $[\text{H}^+]$ in M	pH
Acidic	$[\text{H}^+] > 1.0 \times 10^{-7}\text{M}$	$\text{pH} < 7.00$
Basic	$[\text{H}^+] < 1.0 \times 10^{-7}\text{M}$	$\text{pH} > 7.00$
Neutral	$[\text{H}^+] = 1.0 \times 10^{-7}\text{M}$	$\text{pH} = 7.00$

35. $\text{p}K_w = \text{pH} + \text{pOH} = 14$

36. Dissociation constant or ionization constant (K_a) of a weak acid (HX) in water is given by equation shown below.

$$K_a = \frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]} = \frac{\alpha^2 c}{(1-\alpha)}$$

At a given temperature T, K_a is a measure of the strength of the acid HX i.e., larger the value of K_a , the stronger is the acid.

37. Relation between K_a and $\text{p}K_a$:

$$pK_a = -\log K_a$$

Larger the value of pK_a , weaker is the acid

38. Dissociation constant or ionization constant (K_b) of a weak base (MOH) in water is given by equation shown below.

$$K_b = \frac{[M^+][OH^-]}{[MOH]} = \frac{\alpha^2 c}{(1-\alpha)}$$

At a given temperature T , K_b is a measure of the strength of the base MOH i.e., larger the value of K_b , the stronger is the base.

39. The degree of ionization α is defined as the extent of ionization

40. Relation between K_b and pK_b

$$pK_b = -\log K_b$$

Larger the value of pK_b , weaker is the base.

41. Relation between K_a and K_b

$$K_a \times K_b = K_w$$

42. $pK_a + pK_b = pK_w = 14$ (at 298K)

43. Bronsted acids which can donate more than one protons are called polyprotic acids or polybasic acids.

44. Higher order ionization constants (K_{a2}, K_{a3}) are smaller than the lower order ionization constant (K_{a1}) of a polyprotic acid. The reason for this is that it is more difficult to remove a positively charged proton from a negative ion due to electrostatic forces.

45. Factors affecting acid strength:

- Extent of dissociation of acid (HA) depends on strength and polarity of H-A bond in acid.
- As strength of H-A bond decreases acid strength increases
- As the electronegativity difference between the atoms H and A increases, charge separation between atoms increases, cleavage of the bond becomes easy. Eventually acid dissociates easily increasing acidity.

46. Common ion Effect: If to an ionic equilibrium, a salt containing a common ion is added, the equilibrium shifts in the backward direction. The shift in equilibrium position caused by the addition or presence of an ion involved in the equilibrium reaction is known as common ion effect

47. Interaction of anion or cation of the salt with water to produce an acidic or basic solution is called hydrolysis of salt.

48. Hydrolysis of Salts and the pH of their Solutions

Type of acid and base used to form salt	Example	Type of solution formed by hydrolysis of salt
Salt of strong acid and weak base	NH_4Cl	Acidic solutions
Salt of weak acid and strong base	CH_3COONa	Basic solutions
Salt of weak acid and weak base	$\text{CH}_3\text{COONH}_4$	Neutral solutions
Salt of strong acid and strong base	NaCl	Neutral solutions

49. A buffer solution resists a change in pH caused by dilution or the addition of limited amounts of acid or base.

50. Acidic buffers contain equimolar quantities of a weak acid and one of its salt with a strong base

51. Basic buffers contain equimolar quantities of a weak base and one of its salt with a strong acid

52. Solubility is a measure of amount of solute that can be dissolved in a given amount of solvent at a specific temperature

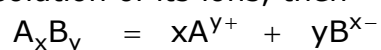
53. In a saturated solution equilibrium exists between the undissolved solute and the solute in the solution.

54. Classification of salts on the basis of solubility

Category I	Soluble	Solubility > 0.1M
Category II	Slightly Soluble	0.01M<Solubility< 0.1M
Category III	Sparingly Soluble	Solubility < 0.01M

55. Solubility product of a salt at a given temperature is equal to product of concentration of the ions in a saturated solution with each concentration term raised to power equal to number of ions produced on dissociation of one mole of substance.

In general if sparingly soluble salt, A_xB_y is in equilibrium with saturated solution of its ions, then



(solid)

Solubility expression is: $K_{sp} = [A^{y+}]^x[B^{x-}]^y$

56. Ionic product is defined as the product of the molar concentrations of the constituent ions, each raised to the power of its stoichiometric coefficient in a any solution.

57. Precipitation will occur only when ionic product exceeds the solubility product

Subject: Chemistry

Class: XI

Chapter: Redox Reactions

Top concepts

1. Redox reactions are those reactions in which oxidation and reduction takes place simultaneously
2. Classical view of redox reactions
 - Oxidation is addition of oxygen / electronegative element to a substance or removal of hydrogen / electropositive element from a substance
 - Reduction is removal of oxygen / electronegative element from a substance or addition of hydrogen / electropositive element to a substance
3. Redox reactions in terms of Electron transfer
 - Oxidation is defined as loss of electrons by any species
 - Reduction is defined as gain of electrons by any species
4. In oxidation reactions there is loss of electrons or increase in positive charge or decrease in negative charge
5. In reduction reactions there is gain of electrons or decrease in positive charge or increase in negative charge
6. Oxidising agents are species which gain one or more electrons and get reduced themselves
7. Reducing agents are the species which lose one or more electrons and gets oxidized themselves
8. Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules. These rules are formulated on the basis that electron in a covalent bond belongs entirely to the more electronegative element.

9. Rules for assigning oxidation number to an atom

- Oxidation number of Hydrogen is always +1 (except in hydrides, it is -1).
- Oxidation number of oxygen in most of compounds is -2. In peroxides it is (-1). In superoxides, it is (-1/2). In OF_2 oxidation number of oxygen is +2. In O_2F_2 oxidation number of oxygen is +1
- Oxidation number of Fluorine is -1 in all its compounds
- For neutral molecules sum of oxidation number of all atoms is equal to zero
- In the free or elementary state, the oxidation number of an atom is always zero. This is irrespective of its allotropic form
- For ions composed of only one atom, the oxidation number is equal to the charge on the ion
- The algebraic sum of the oxidation number of all the atoms in a compound must be zero
- For ions the sum of oxidation number is equal to the charge on the ion
- In a polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must be equal to the charge on the ion

10. Oxidation state and oxidation number are often used interchangeably

11. According to Stock notation the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula

12. Types of Redox Reactions

- **Combination Reactions:** Chemical reactions in which two or more substances (elements or compounds) combine to form a single substance
- **Decomposition Reactions:** Chemical reactions in which a compound break up into two or more simple substances
- **Displacement Reactions:** Reaction in which one ion(or atom)in a compound is replaced by an ion(or atom) of other element
 - a. **Metal Displacement Reactions:** Reactions in which a metal in a compound is displaced by another metal in the uncombined state
 - b. **Non-metal Displacement Reactions:** Such reactions are mainly hydrogen displacement or oxygen displacement reactions
- **Disproportionation Reactions:** Reactions in which an element in one oxidation state is simultaneously oxidized and reduced

13.Steps involved in balancing a Redox reaction by oxidation number method

- Write the skeletal redox reaction for all reactants and products of the reaction
- Indicate the oxidation number of all the atoms in each compound above the symbol of element
- Identify the element/elements which undergo change in oxidation numbers
- Calculate the increase or decrease in oxidation number per atom
- Equate the increase in oxidation number with decrease in oxidation number on the reactant side by multiplying formula of oxidizing agent and reducing agents with suitable coefficients

- Balance the equation with respect to all other atoms except hydrogen and oxygen
- Finally balance hydrogen and oxygen. For balancing oxygen atoms add water molecules to the side deficient in it. Balancing of hydrogen atoms depend upon the medium
 - a. For reactions taking place in acidic solutions add H^+ ions to the side deficient in hydrogen atoms
 - b. For reactions taking place in basic solutions add H_2O molecules to the side deficient in hydrogen atoms and simultaneously add equal number of OH^- ions on the other side of the equation
- Finally balance the equation by cancelling common species present on both sides of the equation

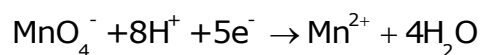
14. Steps involved in balancing a Redox by Ion-Electron Method (Half reaction method)

- Find the elements whose oxidation numbers are changed. Identify the substance that acts as an oxidizing agent and reducing agent
- Separate the complete equation into oxidation half reaction and reduction half reaction
- Balance the half equations by following steps
 - i. Balance all atoms other than H and O
 - ii. Calculate the oxidation number on both sides of equation. Add electrons to whichever side is necessary to make up the difference
 - iii. Balance the half equation so that both sides get the same charge
 - iv. Add water molecules to complete the balancing of the equation
- Add the two balanced equations. Multiply one or both half equations by suitable numbers so that on adding two equations the electrons are balanced

15. Application of Redox reactions: Redox Titrations

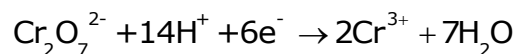
- Potassium permanganate in redox reactions: Potassium permanganate (KMnO_4) is very strong oxidizing agent and is used in determination of many reducing agents like Fe^{2+} , oxalate ions etc. It acts as self indicator in redox reactions.

Equation showing KMnO_4 as an oxidising agent in acidic medium is:

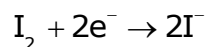


- Acidified Potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) in redox reactions: $\text{K}_2\text{Cr}_2\text{O}_7$ is used as an oxidizing agent in redox reactions. Titrations involving $\text{K}_2\text{Cr}_2\text{O}_7$ uses diphenylamine and potassium ferricyanide (external indicator).

Equation showing $\text{K}_2\text{Cr}_2\text{O}_7$ as an oxidising agent in acidic medium is:



- Iodine (I_2) in redox reactions: I_2 acts as mild oxidising agent in solution according to equation



16. Direct redox reaction: Redox reactions in which reduction and oxidation occurs in same solution (i.e. same reaction vessel). In these reactions transference of electrons is limited to very small distance.

17. Indirect redox reactions: Redox reactions in which oxidation and reduction reactions take place in different reactions vessels and thus transfer of electrons from one species to another does not take place directly

18. Electrochemical cell is a device that converts chemical energy produced in a redox reaction into electrical energy. These cells are also called Galvanic cells or Voltaic cells

19. The electrode at which oxidation occurs is called anode and is negatively charged
20. The electrode at which reduction takes place is called cathode and is positively charged
21. In an electrochemical cell the transfer of electrons takes place from anode to cathode
22. In an electrochemical cell the flow of current is from cathode to anode
23. In the electrochemical cell, the electrical circuit is completed with a salt bridge. Salt bridge also maintains the electrical neutrality of the two half cells
24. A salt bridge is a U shaped tube filled with solution of inert electrolyte like sodium chloride or sodium sulphate which will not interfere in the redox reaction. The ions are set in a gel or agar agar so that only ions flow when inverted
25. Electrical potential difference developed between the metal and its solution is called electrode potential. It can also be defined as tendency of an electrode in a half cell to gain or lose electrons
26. Oxidation potential is the tendency of an electrode to lose electrons or to get oxidized

27.Reduction potential is the tendency of an electrode to gain electrons or get reduced

28.In an electrochemical cell, by the present convention, the electrode potentials are represented as reduction potential

29.The electrode having a higher reduction potential will have a higher tendency to gain electrons

30. By convention, the standard electrode potential of hydrogen electrode is 0.00 volts

31. A redox couple is defined as having together oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction

32.The difference between the electrode potentials of eth two electrodes constituting the electrochemical cell is called EMF(Electromotive force) or the cell potential

$$33.EMF = E^{\ominus}_{\text{cathode}} - E^{\ominus}_{\text{anode}}$$

34. A negative E^{\ominus} means that the redox couple is a stronger reducing agent than the H^+/H_2 couple

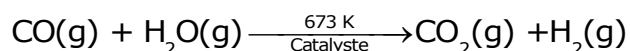
35.A positive E^{\ominus} means that the redox couple is a weaker reducing agent than the H^+/H_2 couple

Subject: Chemistry
Class: XI
Chapter: Hydrogen
Top concepts

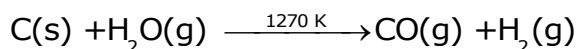
1. Hydrogen is the lightest element known since it has an atomic mass of 1.0079
2. Hydrogen resembles Group 1 elements because it has electronic configuration ($1s^1$) similar to alkali metals and can lose one electron to form unipositive ions
3. Hydrogen also resembles elements of group 17 (Halogens) because hydrogen just like halogens needs just one electron to acquire the configuration of the nearest noble gas i.e. Helium
4. Hydrogen has three isotopes - Protium, deuterium and tritium

Isotopes	No. of protons(p)	No. of neutrons(n)	No. of electrons(e)	Mass number (M)
Protium (${}^1_1\text{H}$)	1	0	1	1
Deuterium(${}^2_1\text{H}$ or D)	1	1	1	2
Tritium ${}^3_1\text{H}$ or D	1	2	1	3

5. In elemental form Hydrogen exists as a diatomic molecule H_2 and is called dihydrogen
6. Water gas is name given to the mixture of carbon monoxide and hydrogen. Water gas is also called synthesis gas or syngas
7. The production of dihydrogen can be increased by reacting carbon monoxide of syngas mixtures with steam in the presence of iron chromate as catalyst. This reaction is called as water gas shift reaction



8. The process of producing 'syngas' from coal is called 'coal gasification'



9. Physical properties of dihydrogen
 - Hydrogen is a colorless, odorless and a tasteless gas
 - Hydrogen is highly combustible

- Hydrogen is lighter than air
- Hydrogen is insoluble in water

10. The binary compounds of hydrogen with other elements are called hydrides

11. Ionic hydrides or Saline hydrides are binary compounds of hydrogen with s- block elements which are highly electropositive

12. Covalent or molecular hydrides are binary compounds of hydrogen with elements of comparatively high electronegativity such as p- block elements

13. Covalent hydrides are classified as electron rich, electron deficient or electron precise hydrides

- Electron rich hydrides have excess electrons than required to form normal covalent bonds. Excess electrons are present as lone pairs .Examples of this are hydrides of group 15, 16 and 17 such as ammonia, water etc.
- Electron deficient hydrides do not have sufficient number of electrons to form normal covalent bonds. Examples of this are hydrides of group 13
- Electron precise hydrides have required number of electrons to write the conventional Lewis structure. Elements of group 14 like carbon, silicon etc. form these types of hydrides

14. Metallic hydrides or Interstitial hydrides are the compounds of hydrogen with transition metals of group 3,4,5 of d-block, Cr metal of group 6, f-block element

15. Since the composition of metallic hydrides does not correspond to simple whole number ratio they are also called non-stoichiometric hydrides

16. Due to high polarity, water dissolves most of the polar substances

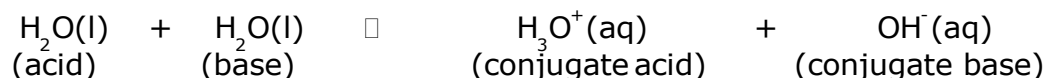
17. Due to hydrogen bonding water has high boiling point and high heat of vaporization

18. Density of water is more than that of ice

19. Water has maximum density (1g/cm^3) at $4\text{ }^\circ\text{C}$

20. Water is amphoteric in nature because it has the ability to act as an acid as well as base

21. Autoprotolysis or self ionization of water: One water molecule acts as an acid by donating a proton to another water molecule which acts as a base. This is also called autoprotolysis of water



22. Water free from soluble salts of calcium and magnesium is called soft water

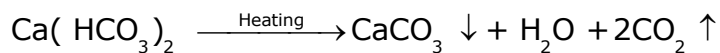
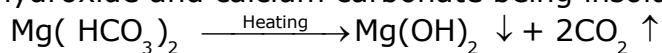
23. Water containing soluble salts of calcium and magnesium in form of hydrogen carbonate, chlorides and sulphates is called hard water

24. Temporary hardness in water is due to soluble salts of hydrogen carbonates of magnesium and calcium

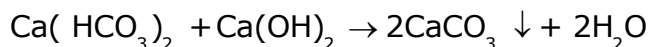
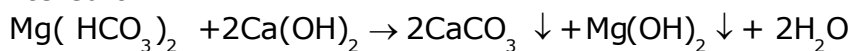
25. Permanent hardness in water is due to soluble salts of chlorides and sulphates of calcium and magnesium

26. Methods to remove temporary hardness

- Boiling: During boiling soluble magnesium hydrogen carbonate is changed to magnesium hydroxide and calcium hydrogen carbonate on heating gets converted into calcium carbonate. Magnesium hydroxide and calcium carbonate being insoluble is filtered off

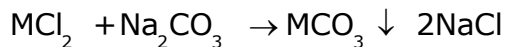


- Clark's method: Calculated amount of calcium hydroxide is added to a given amount of water. It precipitates out calcium as calcium carbonate and magnesium as magnesium hydroxide which can be filtered off

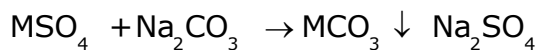


27. Methods to remove permanent hardness

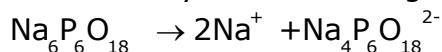
- Treatment with washing soda (sodium carbonate) : Washing soda reacts with hard water forming insoluble metal carbonate which can be filtered and removed



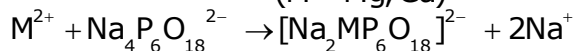
(M = Mg, Ca)



- Calgon's method: Calcium and magnesium ions are rendered ineffective by addition of sodium hexa metaphosphate which is commercially known as Calgon

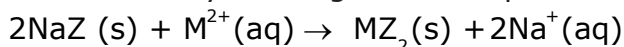


(M = Mg, Ca)

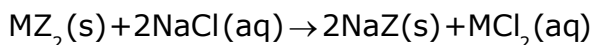


The complex anion keeps the Mg^{2+} and Ca^{2+} ions in solution

- Ion-exchange method: In this method ions responsible for hardness of water are exchanged by certain less damaging ions present in some chemical compound like sodium aluminium silicate ($NaAlSiO_4$) which can also be written as NaZ (also called as zeolites). When this is added in hard water, exchange reactions take place. Zeolite is said to be exhausted when all the sodium in it is used up. It is regenerated for further use by treating with an aqueous sodium chloride solution



(M = Mg, Ca)



- Synthetic resins method: These are insoluble polymeric solids having giant hydrocarbon network containing reactive acidic or basic groups. They perform function similar to zeolites but they are superior to zeolites because they can remove all types of ion in water.

28. Hydrogen peroxide can act as oxidizing agent as well as reducing agent in both acidic and alkaline medium

29. Commercially, hydrogen peroxide is generally sold as its percentage solution. For example 10 volume of hydrogen peroxide means that 1 liter of this hydrogen peroxide solution will give 10 liter of oxygen at STP

30. Hydrogen peroxide decomposes to form water and oxygen, on exposure to sunlight. Hence it is stored in wax-lined glass or plastic vessel

31. D_2O is called heavy water. It is manufactured by electrolytic enrichment of normal water. It is used as moderator in nuclear reactor.

Subject: Chemistry
Class: XI
Chapter: The s-Block Elements
Top concepts

1. The s-block elements of the periodic table are those in which the last electron enters the outermost s-orbital
2. Elements of group 1 are called alkali metals and elements of group 2 are called alkaline earth metal
3. Group 1 elements: Alkali metals
 - (i) Group 1 elements are called alkali metals because they form hydroxides on reaction with water which are strongly alkaline in nature
 - (ii) The general electronic for alkali metals is [noble gas] ns^1

Electronic configuration of alkali metals:

Element	Symbol	Atomic Number	Electronic configuration
Lithium	Li	3	[He] $2s^1$
Sodium	Na	11	[Ne] $3s^1$
Potassium	K	19	[Ar] $4s^1$
Rubidium	Rb	37	[Kr] $5s^1$
Caesium	Cs	55	[Xe] $6s^1$
Francium	Fr	87	[Rn] $7s^1$

(iii) Trends in properties of alkali metals:

1. Atomic and ionic radii: Alkali metals have the largest atomic and ionic radii in their respective periods of the periodic table. On moving down the group, the atomic and ionic radii increase

Explanation: As we move in a period, the atomic radius and ionic radius tend to decrease due to increase in the effective nuclear charge. Therefore alkali metals have largest atomic and ionic radii in their respective group. On moving down the group there is increase in the number of shells. Thus there is an increase in distance between nucleus and outermost electrons which eventually increases atomic and ionic radii

2. Ionisation enthalpies: Alkali metals have the lowest ionization enthalpy in each period. Within the group, the ionization enthalpies of alkali metals decrease down the group

Explanation: Atoms of alkali metals are largest in their respective periods and therefore, the valence electrons are loosely held by the nucleus. By losing the valence electron, they acquire stable noble gas configuration. This accounts for their ease to lose electrons and hence they, have low ionization enthalpies

3) The second ionization enthalpies of alkali metals are very high.

Explanation: When an electron is removed from the alkali metals, they form monovalent cations which have very stable electronic configurations (same as that of noble gases). Therefore it becomes difficult to remove the second electron from the stable noble gas configurations and hence their second ionization enthalpy values (IE_2) are very high

4. Melting and boiling points: Alkali metals are soft and have low melting and boiling points

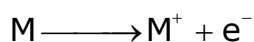
Explanation: Alkali metals have only one valence electron per metal atom and therefore, the energy binding the atoms in the crystal lattice of the metal is low. Thus, the metallic bonds in these metals are not very strong and consequently, their melting and boiling points decrease on moving down from Li to Cs

5. Density: Densities of alkali metals are quite low as compared to other metals. The densities increase on moving down the group. But K is lighter than Na

Explanation: The densities of metallic elements depend upon the type of packing of atoms in metallic state and also on their size. The alkali metals have close packing of metal atoms in their lattice and because of the large size of their atoms, they have low densities. On moving down the group from Li to Cs, there is increase in atomic size as well as atomic mass. But the increase in atomic mass is more and compensates the increase in atomic size. As a result, the densities (mass/volume) of alkali metals gradually increase from Li to Cs. K is lighter than Na due to increase in atomic size of K

6. Electropositive or metallic character: All the alkali metals are strongly electropositive or metallic in character

Explanation: The electropositive character of an element is expressed in terms of the tendency of its atom to release electrons:



As alkali metals have low ionization enthalpies, their atoms readily lose their valence electron. These elements are, therefore, said to have strong electropositive or metallic character. Since, the ionization energies decrease down the group, the electron releasing tendency or electropositive character is expected to increase down the group

7. Oxidation states: All the alkali metals predominantly exhibit an oxidation state of +1 in their compounds

Explanation: Alkali metals have only one electron in their valence shell and therefore can lose the single valence electron readily to acquire the stable configuration of a noble gas. Since the second ionization energies are very high, they cannot form divalent ions. Thus, alkali metals are univalent and form ionic compounds

8. Characteristic flame colouration: All the alkali metals and their salts impart characteristic flame colouration

Explanation: Alkali metals have very low ionization enthalpies. The energy from the flame of Bunsen burner is sufficient to excite the electrons of alkali metals to higher energy levels. Excited state is quite unstable and therefore when these excited electrons come back to their original energy levels, they emit extra energy, which falls in the visible region of the electromagnetic spectrum and thus appear coloured. Characteristic flame colouration by different alkali metals can be explained on the basis of difference in amount of energy absorbed for excitation of the valence electron

9. Photoelectric effect: Phenomenon of ejection of electrons when electromagnetic radiation of suitable frequency strikes metal surface is called photoelectric effect. Alkali metals exhibit photoelectric effect

Explanation: Alkali metals have low ionization enthalpies and therefore, the electrons are easily ejected when exposed to light. Among alkali metals, caesium has lowest ionization enthalpy and hence it shows photoelectric effect to the maximum extent

10. Hydration of alkali metal ions: Alkali metal ions are highly hydrated.

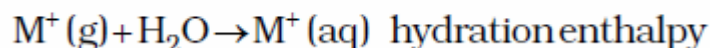
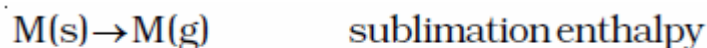
Explanation: Smaller the size of the ion, the greater is the degree of hydration. Thus, Li^+ ion gets much more hydrated than Na^+ ion which is more hydrated than K^+ ion and so on. Therefore, the extent of hydration decreases from Li^+ to Cs^+ . As a result of larger hydration of Li^+ ion than Na^+ ion, the effective size of hydrated Li^+ ion is more than that of hydrated Na^+ ion.

Hydrated ionic radii decrease in the order: $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$

Due to extensive hydration, Li^+ ion has lowest mobility in water.

11. Reducing Nature: Alkali metals are strong reducing agents. This is due to their greater ease to lose electrons

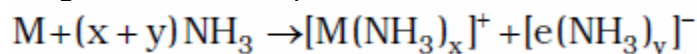
Explanation: Lithium is the strongest reducing agent. Tendency to act as reducing agent depends on energy requirement involved in three processes i.e. sublimation, ionization and hydration.



Lithium being small in size has high ionization enthalpy. On the other hand because of small size it is extensively hydrated and has very high hydration enthalpy. This high hydration enthalpy compensates the high energy needed to remove electron (in second step). Thus Li has greater tendency to lose electrons in solution than other alkali metals. Thus Li is the strongest reducing agent.

12. Solutions in liquid ammonia: Alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature

Explanation: In solution the alkali metal atom readily loses the valence electron. Both the cation and the electron combine with ammonia to form ammoniated cation and ammoniated electron. The blue colour of the solution is due to the ammoniated electron which absorbs energy in the visible region of light and thus imparts blue colour to the solution



13. Basic strength of hydroxides: Hydroxides of alkali metals are strongly basic and basic strength increases down the group

Explanation: The Metal—OH bond in the hydroxides of alkali metals is very weak and it can easily ionize to form M^+ and OH^- ions. This accounts for their basic character. Since the ionization energy decreases down the group, the bond between metal and oxygen becomes weak. Therefore, the basic strength of the hydroxides increases accordingly.

(iv) Important Compounds of alkali metals:

Name of compound	Name of process & Brief about the process	Related chemical equations
Sodium Carbonate (Washing Soda), $Na_2CO_3 \cdot 10H_2O$	Solvay Process: When CO_2 gas is passed through a brine solution saturated with ammonia, sodium bicarbonate is formed. Sodium bicarbonate on heating forms sodium carbonate	$2NH_3 + H_2O + CO_2 \rightarrow (NH_4)_2CO_3$ $(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$ $NH_4HCO_3 + NaCl \rightarrow NH_4Cl + NaHCO_3$ $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
Sodium Chloride, NaCl	Evaporation of sea water: Evaporation of sea water gives crude salt which contains impurities of $CaSO_4$, Na_2SO_4 etc. To obtain pure sodium chloride,	-

	the crude salt is dissolved in minimum amount of water and filtered to remove insoluble impurities. The solution is then saturated with HCl gas. Crystals of pure sodium chloride separate out.	
Sodium Hydroxide (Caustic Soda), NaOH	Electrolysis of NaCl in Castner- Kellner cell: A brine solution is electrolysed using a mercury cathode and a carbon anode. Na metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode. Sodium amalgam on treatment with water forms sodium hydroxide	$\text{Cathode : } \text{Na}^+ + \text{e}^- \xrightarrow{\text{Hg}} \text{Na - amalgam}$ $\text{Anode : } \text{Cl}^- \rightarrow \frac{1}{2} \text{Cl}_2 + \text{e}^-$ $2\text{Na-amalgam} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + 2\text{Hg} + \text{H}_2$
Sodium Hydrogencarbonate (Baking Soda), NaHCO ₃	NaHCO ₃ is made by saturating a solution of sodium carbonate with carbon dioxide	$\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow 2\text{NaHCO}_3$

4. Group 2 elements: Alkaline earth metals

(i) Group 2 elements are called alkaline earth metals because their oxides and hydroxides are alkaline in nature and these metal oxides are found in the earth's crust

(ii) The general electronic configuration for alkaline earth metals is [noble gas] ns²

Electronic configuration of alkaline earth metals:

Element	Symbol	Atomic Number	Electronic configuration
Beryllium	Be	4	[He]2s ²
Magnesium	Mg	12	[Ne]3s ²
Calcium	Ca	20	[Ar]4s ²
Strontium	Sr	38	[Kr]5s ²
Barium	Ba	56	[Xe]6s ²

Radium	Ra	88	$[\text{Rn}]7s^2$
--------	----	----	-------------------

(iii) Trends in properties of alkaline metals and their comparison with alkali metals:

1. Atomic and ionic radii: Atomic and ionic radii of alkaline earth metals increases down the group and are smaller than the corresponding members of the alkali metals

Explanation: Alkaline earth metals have a higher nuclear charge and therefore, the electrons are attracted more towards the nucleus. As a result, their atomic and ionic radii are smaller than those of alkali metals. Atomic and ionic radii increase on moving down the group, due to screening effect and increases in the number of the shells.

2. Ionisation enthalpies: Alkaline earth metals have low ionization enthalpies due to their large size

Explanation: Although IE_1 values of alkaline earth metals are higher than those of alkali metals, the IE_2 values of alkaline earth metals are much smaller than those of alkali metals

In case of alkali metals (for e.g. Na) the second electron is to be removed from a cation which has already acquired a noble gas configuration. However in the alkaline earth metals (for e.g. Mg), the second electron is to be removed from a monovalent cation, i.e. Mg^+ : $(1s^2 2s^2 2p^6 3s^1)$ which still has one electron in the outermost shell. Thus, the second electron in Mg can be removed easily.

3. Melting and boiling points: Alkaline earth metals have higher melting and boiling points than the corresponding alkali metals

Explanation: Atoms of alkaline earth metals have smaller size as compared to alkali metals. Due to two valence electrons in alkaline earth metals, metallic bond is stronger than alkali metals. Hence alkaline earth metals have higher melting and boiling points than corresponding alkali metals

4. Electropositive or metallic character: The electropositive character increases down the group i.e., from Be to Ba but alkaline earth metals are not as strongly electropositive as the alkali metals

Explanation: Because of the low ionization enthalpies of alkaline earth metals, they are strongly electropositive in nature. However, these are not as strongly electropositive as the alkali metals of group 1 because of comparatively higher ionization energies

5. Oxidation states: All the alkaline earth metal exhibits an oxidation state of +2 in their compounds

Explanation: Alkaline earth metals exhibit oxidation state of +2 as they can lose two electrons and form bivalent ions. Following are the reason for the easy removal of two electrons

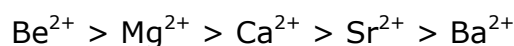
(i) Divalent cations have stable noble gas configuration

(ii) In solution, the +2 ions of alkaline earth metals are extensively hydrated and the high hydration energies of E^{2+} ions make them more stable than E^+ ions. It is observed that the amount of energy released when E^{2+} ion is dissolved in water is much more than that for E^+ ions. This large amount of extra energy released in the hydration of +2 ions is more and compensates the second ionization enthalpy required for the formation of such E^{2+} ions

(iii) In the solid state, the divalent cations form stronger lattices than monovalent cations and therefore, a lot of energy called lattice enthalpy is released. It is the greater lattice enthalpy of E^{2+} ion which compensates for the high second ionization enthalpy and is responsible for its greater stability as compared to E^+ ion

6. Hydration enthalpy: Hydration enthalpies of alkaline earth metal ions decrease down the group and are larger than those of the corresponding alkali metal ions

Explanation: This is due to smaller size of alkaline earth metal ions as compared to corresponding alkali metal ions of the same period. Due to smaller size, the alkaline earth metal ions are strongly attracted by polar water molecules and hence have larger hydration enthalpies. The hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group.



7. Basic strength of hydroxides: Hydroxides of alkaline earth metals are less basic than alkali metals of the corresponding period.

Explanation: Lesser basic strength of hydroxides of alkaline earth metal hydroxides is due to their high ionization enthalpies, small ionic size and dipositive charge on the ions

As a result, the Metal—O bond in these hydroxides is relatively stronger than that of corresponding alkali metals and therefore, does not break easily. Therefore, they are less basic than corresponding alkali metals.

(iii) Important Compounds of alkaline earth metals:

Name of compound	Name of process & Brief about the process	Related chemical equations
Calcium Oxide or Quick Lime, CaO	It is prepared by heating limestone (CaCO_3) in a rotary kiln at 1070-1270 K. The carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion	$\text{CaCO}_3 \xrightleftharpoons{\text{heat}} \text{CaO} + \text{CO}_2$

Calcium Hydroxide (Slaked lime), Ca(OH)_2	It is prepared by adding water to quicklime, CaO	$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
Calcium Carbonate, CaCO_3	It is prepared by passing carbon dioxide through slaked lime. Excess of carbon dioxide should be avoided since this leads to the formation of water soluble calcium hydrogencarbonate	$\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$
	It is prepared by the addition of sodium carbonate to calcium chloride.	$\text{CaCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow \text{CaCO}_3 + 2\text{NaCl}$
Calcium Sulphate (Plaster of Paris), $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$	It is obtained when gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, is heated to 393 K	$2(\text{CaSO}_4 \cdot 2\text{H}_2\text{O}) \rightarrow 2(\text{CaSO}_4) \cdot \text{H}_2\text{O} + 3\text{H}_2\text{O}$

5. Diagonal relationship: It is observed that some elements of second period show similarities with the elements of third period present diagonally to each other though belonging to different groups. This is called diagonal relationship

Cause of diagonal relationship: Reason is the similarity in properties such as electronegativity, ionization enthalpy size, charge/radius ratio, etc. between the diagonal elements. On moving from left to right across a period, for example, the electronegativity increases and while moving down group electronegativity decreases. Therefore, on moving diagonally, the two opposing tendencies almost cancel out and the electronegativity values remain almost same as we move diagonally. Thus, the diagonal pairs have many similar properties.

Following pairs exhibit diagonal similarity:

- (i) Lithium – Magnesium
- (ii) Beryllium – Aluminium
- (iii) Boron – Silicon

6. Some important reasons:

(1) Lithium carbonate decomposes at a lower temperature whereas sodium carbonate decomposes at a higher temperature

Explanation: Lithium carbonate decomposes at a lower temperature whereas sodium carbonate decomposes at a higher temperature because lithium carbonate is unstable to heat. Lithium ion is smaller in size than sodium ion and thus polarizes the carbonate ion to a greater extent than sodium ion leading to the formation of more stable lithium oxide and carbon dioxide.

(2) Potassium carbonate cannot be prepared by Solvay process

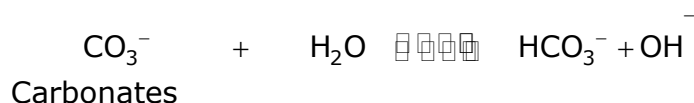
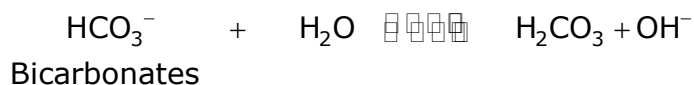
Explanation: Unlike NaHCO_3 , the intermediate KHCO_3 formed during reaction in Solvay process, is highly soluble in water and thus cannot be taken out from solution to obtain K_2CO_3 . Hence, K_2CO_3 cannot be prepared by Solvay process.

(3) Alkali and alkaline earth metals cannot be obtained by chemical reduction methods

Explanation: Alkali and alkaline earth metals are themselves very strong reducing agents and therefore cannot be reduced by chemical reduction methods

(5) An aqueous solution of carbonates and bicarbonates of alkali metals is alkaline

Explanation: The aqueous solution of carbonates and bicarbonates of alkali metals is alkaline due to the hydrolysis of these salts which produces hydroxide ions.



(5) LiCl is more covalent than KCl

Explanation: According to Fajan's rule smaller the size of cation and larger the size of anion greater is the covalent character of ionic bond. Li is small in size than K, thus Li^+ has a high charge density. Thus polarizing power of Li^+ is higher than K^+ , hence LiCl is more covalent than KCl.

(6) BaO is soluble but BaSO_4 is insoluble in water

Explanation: Size of O^{2-} ion is smaller than SO_4^{2-} . Since a bigger anions stabilizes bigger cation more than a smaller cation stabilizes a bigger anion,

lattice enthalpy of BaO is smaller than BaSO₄. BaO is soluble as hydration energy is more than lattice energy but BaSO₄ (as hydration energy is less than lattice energy) is insoluble in water.

(7) Solubility of alkaline earth metal carbonates and sulphates in water decrease down the group

Explanation: The size of anions being much larger compared to cations, the lattice enthalpy will remain almost constant within a particular group. Since the hydration enthalpies decrease down the group, solubility will decrease as found for alkaline earth metal carbonates and sulphates.

(8) Alkali metals are not found in nature

Explanation: Alkali metals are highly reactive because of low ionization enthalpy value and therefore are not found in nature. They are present in combined state only in form of halides, oxides etc.

(9) Sodium is less reactive than potassium

Explanation: Ionization Energy of potassium is less than sodium because of large size or less effective nuclear charge. Thus, potassium is more reactive than sodium

(10) NaOH is a stronger base than LiOH

Explanation: The M-OH bond in hydroxides of alkali metal is very weak and can easily ionize to form M⁺ ions and OH⁻ ions. This accounts for the basic character. Since ionization energy decreases down the group bond between metal and oxygen becomes weak. Therefore basic strength of hydroxides increases accordingly. Thus NaOH is a stronger base than LiOH

(11) Alkali metals are kept in paraffin or kerosene

Explanation: Alkali metals react explosively with water forming metal hydroxides along with hydrogen. Hydrogen gas released immediately catches fire. Thus alkali metals are highly sensitive towards air and water and hence are kept therefore in kerosene or paraffin oil

(12) Except for Be and Mg, the alkaline earth metal salts impart characteristic colours to the flame

Explanation: Beryllium and magnesium atoms are comparatively smaller and their ionization energies are very high. Hence, the energy of the flame is not sufficient to excite their electrons to higher energy levels. These elements, therefore, do not give any colour in Bunsen flame

Subject: Chemistry
Class: XI
Chapter: The p-Block Elements
Top concepts

1. Elements in which the last electron enters the p-subshell of their outermost energy level are called p-block elements. The elements belonging to groups 13 to 18 constitute p-block elements.

2. General electronic configuration for elements belonging to p-block is ns^2np^{1-6}

3. General characteristics of p-block elements

1. Variation in oxidation states:

Group	13	14	15	16	17	18
General electronic configuration	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
First member of group	B	C	N	O	F	He
Group oxidation state	+3	+4	+5	+6	+7	+8
Other oxidation states	+1	+2, -4	-3,+3	-2,+2,+4,	+5,+3,-1,+1	+2,+4,+6

2. Metallic and non-metallic character: Elements which show characteristics of both metals and non-metals and are called metalloids. Common metalloids in p-block elements are Si, Ge, As, Sb, Te, Po, At.

3. Differences in behaviour of first element of each group:

(a) Covalence upto four: First member of each group belongs to second period elements and have only four valence orbitals i.e., one 2s and three 2p orbitals. They do not have vacant d-orbitals in their valence shell. Therefore each member has maximum covalence of four (using one 2s and three 2p orbitals). Next members belonging to third or higher periods have vacant d-orbitals. For example elements of third period of p-block with the electronic configuration $3s^2 3p^x$ has vacant 3d-orbitals lying between 3p and 4s energy levels. Therefore, they can easily expand their octets and can show covalence above four. For example:

(i) Boron forms only BF_4^- (coordination number of B is four) whereas aluminium forms AlF_6^{3-} (coordination number of Al is six).

(ii) Carbon can form only tetrahalides (CX_4 , X = F, Cl, Br, I) whereas other members can form hexahalides, SF_6 , $SiCl_6^{2-}$ etc.

(iii) Nitrogen forms only NF_3 while phosphorous forms pentahalides like PCl_5 , PF_5 etc.

(iv) Fluorine does not form FCl_3 (F has more than octet) while chlorine forms ClF_3 (Cl has extended octet)

(b) Reactivity: Elements of second period do not have d – orbitals and are less reactive as compared to elements of third period which have d – orbitals. For example tetrahalides of carbon are not hydrolysed by water whereas tetrahalides of other elements of group 14 are readily hydrolysed (for e.g. Si)

(c) Tendency to form multiple bonds: Because of combined effect of smaller size and availability of p – orbitals, the first member of each group shows greater tendency to form $p\pi - p\pi$ multiple bonds either with itself (such as $\text{C} = \text{C}$, $\text{C} \equiv \text{C}$, $\text{N} \equiv \text{N}$, $\text{O} = \text{O}$) or with other members of the second period of elements (such as $\text{C} = \text{O}$, $\text{C} \equiv \text{N}$, $\text{N} = \text{O}$, etc.). Heavier elements of the group do not have strong tendency to form this type of $p\pi - p\pi$ bonding. The heavier elements do form π - bonding but they involve d – orbitals and form $d\pi - p\pi$ or $d\pi - d\pi$ bonding.

5. Group 13 Elements: The Boron family

1. Electronic Configurations: Elements of group 13 have the general outer electronic configuration of ns^2np^1

Elements	Symbol	Atomic No.	Electronic Configuration
Boron	B	5	$[\text{He}] 2s^2 2p^1$
Aluminium	Al	13	$[\text{Ne}] 3s^2 3p^1$
Gallium	Ga	31	$[\text{Ar}] 3d^{10} 4s^2 4p^1$
Indium	In	49	$[\text{Kr}] 4d^{10} 5s^2 5p^1$
Thallium	Tl	81	$[\text{Xe}] 4f^{14} 5d^{10} 6s^2 6p^1$

2. Trends in properties of group 13 elements:

No.	Property	Variation along the group
1	Atomic radii	Atomic radii increases down the group but not regularly
2	Ionization Enthalpy	Ionization enthalpy decreases down the group but not regularly
3	Electronegativity	Down the group, electronegativity first decreases from B to Al and then increases marginally

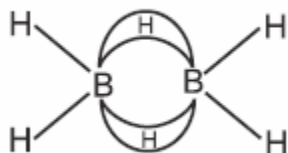
3. Some important compounds of boron:

No.	Name of compound	Formula of compound
1	Borax	$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$

2	Orthoboric acid	H_3BO_3
3	Diborane	B_2H_6

4. Structure of diborane (B_2H_6):

The electron diffraction studies have shown bridged structure for diborane as shown below



Four hydrogen atoms (2 on the left and 2 on the right) are known as terminal hydrogens. These are different than the other two hydrogen atoms, which are known as bridged hydrogens. The two boron atoms and the four terminal H-atoms lie in the same plane while the two bridging H-atoms are in a plane perpendicular to the rest of the molecule. There are two types of bonds in the molecule:

- (i) The four terminal hydrogens are bonded by normal covalent bonds formed by sharing of one electron each from B and H atoms
- (ii) The bridged hydrogens are bonded by three centre electron pair bond, which involve one electron pair only (i.e. two electrons) but binds three atoms i.e. two B atoms and one H atom.

Boron atom undergoes sp^3 hybridisation involving 2s and all the three 2p orbitals including one empty orbital. The four sp^3 hybrid orbitals adopt tetrahedral arrangement. Two hybrid orbitals of each B atom overlap with 1s orbital of two H atoms. Of the hybrid orbitals left on each B atom, one contains an unpaired electron while the other is vacant. The hybrid orbital containing an unpaired electron of one B atom and the vacant hybrid orbital of the second B atom overlaps simultaneously with 1s orbitals of H atom to form B-H-B bridge bond. B-H-B bond is called three centre electron pair bond.

5. Group 14 Elements: Carbon Family

1. Electronic Configuration: Elements of group 14 have the general outer electronic configuration of ns^2np^2

Element	Symbol	Atomic No.	Electronic configuration
Carbon	C	6	[He] $2s^2 2p^2$
Silicon	Si	14	[Ne] $3s^2 3p^2$
Germanium	Ge	32	[Ar] $3d^{10} 4s^2 4p^2$
Tin	Sn	50	[Kr] $4d^{10} 5s^2 5p^2$

Lead	Pb	82	[Xe] 4f ¹⁴ 5d ¹⁰ 6s ² 6p ²
------	----	----	--

2. Trends in properties of group 14 elements:

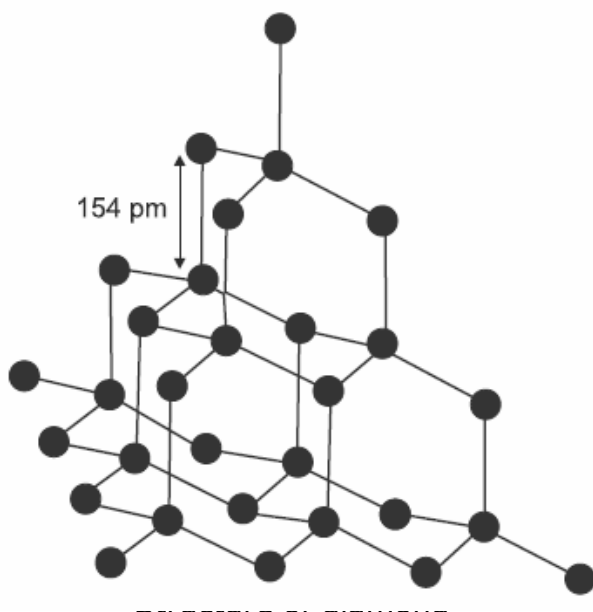
No.	Property	Variation along the group
1	Atomic radii	An atomic radius increases down the group. There is a considerable increase in covalent radius from C to Si, thereafter from Si to Pb a small increase in radius is observed
2	Ionization Enthalpy	Ionization enthalpy decreases down the group but not regularly. Small decrease in $\Delta_i H$ from Si to Ge to Sn and slight increase in $\Delta_i H$ from Sn to Pb is the consequence of poor shielding effect of intervening <i>d</i> and <i>f</i> orbitals and increase in size of the atom
3	Electronegativity	Due to small size, the elements of this group are slightly more electronegative than group 13 elements. The electronegativity values for elements from Si to Pb are almost the same

(3) Allotropes of Carbon: Allotrope forms or allotropes are the different forms of the same elements having different physical properties but similar chemical properties.

1. Crystalline form: Diamond and graphite are two crystalline forms of carbon having well defined structure
2. Amorphous form: Coal, wood charcoal, animal charcoal, lamp black, coke, etc. are many amorphous forms of carbon such as

(5) Structure of different allotropes of carbon:

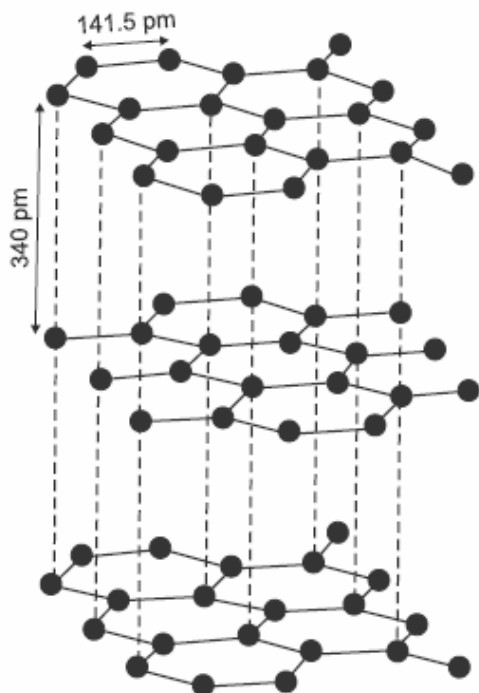
(a) Diamond: Diamond has a network structure consisting of a very large number of carbon atoms bonded to each other. Each carbon atom is sp^3 hybridized and is bonded to four other carbon atoms by single covalent bonds. There is a three – dimensional network of strong covalent bonds in diamond. This makes diamond an extremely hard crystal with very high melting point (about 3843 K).



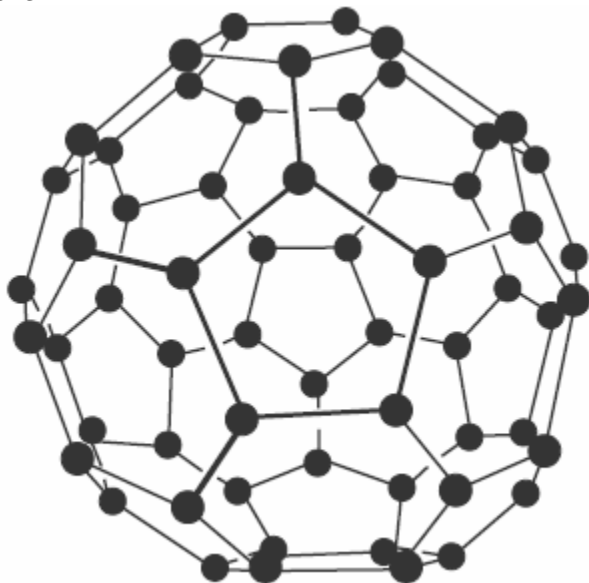
Since all the valence electrons of carbon are strongly held in carbon – carbon bonds, diamond is a poor conductor of electricity.

(b) Graphite: In graphite, each carbon atom undergoes sp^2 hybridisation and is covalently bonded to three other carbon atoms by single bonds. The fourth electron on each carbon atom forms π bonds. In this way, graphite consists of hexagonal rings in two dimensions. The C–C covalent distance in rings is 141.5 pm indicating strong bonding. These arrays of ring form layers. The layers are separated by a distance of 340 pm. The large distance between these layers indicates that only weak van der Waals' forces hold these layers together. The weak van der Waals' forces which hold these layers together are responsible for soft nature of graphite.

Because of the weak Van der Waals' forces between layers, one layer can move over the other layer and this account for the slippery nature of graphite. Therefore, graphite is used as lubricant. The electrons forming π bonds in the rings of graphite are delocalized and are relatively free to move under the influence of electric field. Therefore, graphite is a good conductor of electricity.



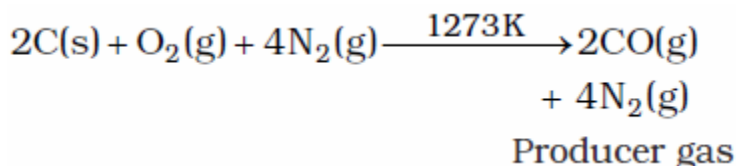
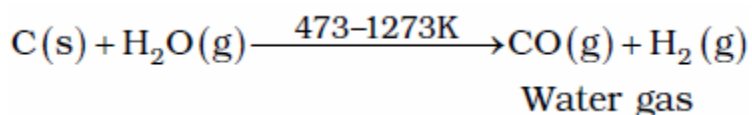
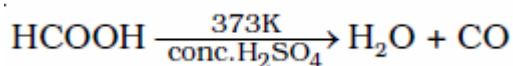
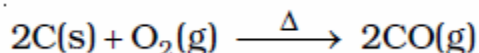
(c) Fullerene: It contains 20 six-membered rings and 12 five-membered rings. All the carbon atoms are equal and they undergo sp^2 hybridisation. Each carbon atom forms three sigma bonds with other three carbon atoms. The remaining electron at each carbon is delocalised in molecular orbitals, which in turn give aromatic character to molecule. This ball shaped molecule has 60 vertices and each one is occupied by one carbon atom and it also contains both single and double bonds with C-C distances of 143.5 pm and 138.3 pm respectively. Spherical fullerenes are also called bucky balls in short.



4. Some important Compounds of carbon and silicon

(i) Carbon Monoxide (CO)

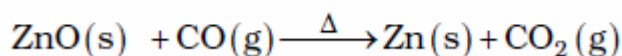
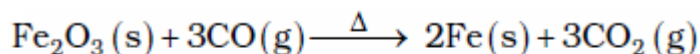
(a) Preparation of carbon monoxide (CO):



(b) Properties of carbon monoxide (CO):

(i) It is colourless, odourless and almost water insoluble gas

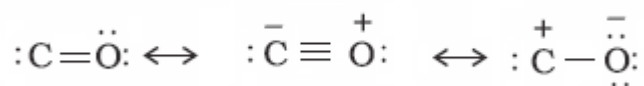
(ii) It is a powerful reducing agent and reduces almost all metal oxides other than those of the alkali and alkaline earth metals, aluminium and a few transition metals



(iii) CO is highly poisonous because it has ability to form a complex with haemoglobin, which is about 300 times more stable than the oxygen-haemoglobin complex. This prevents haemoglobin in the red blood corpuscles from carrying oxygen round the body and ultimately resulting in death.

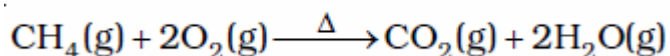
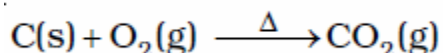
(c) Structure of carbon monoxide (CO): In CO molecule, there are one σ and two π bonds between carbon and oxygen: $\text{C} \equiv \text{O}:$. Because of the presence of a lone pair on carbon, CO molecule acts as a donor and reacts with certain metals when heated to form metal carbonyls.

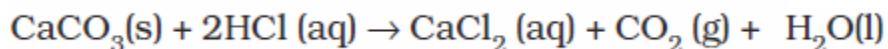
Carbon monoxide is regarded as resonance hybrid of the following structures:



(ii) Carbon dioxide (CO₂)

(a) Preparation of carbon dioxide (CO₂) :

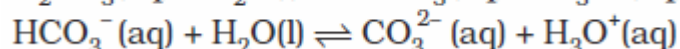
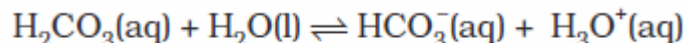




(b) Properties of carbon dioxide (CO₂):

(i) It is a colourless and odourless gas

(ii) With water, it forms carbonic acid, H₂CO₃ which is a weak dibasic acid and dissociates in two steps:

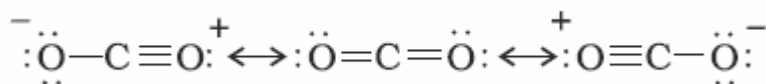


H₂CO₃/HCO₃⁻ buffer system helps to maintain pH of blood between 7.26 to 7.42

(iii) Being heavy and non-supporter of combustion carbon dioxide is used as fire extinguisher

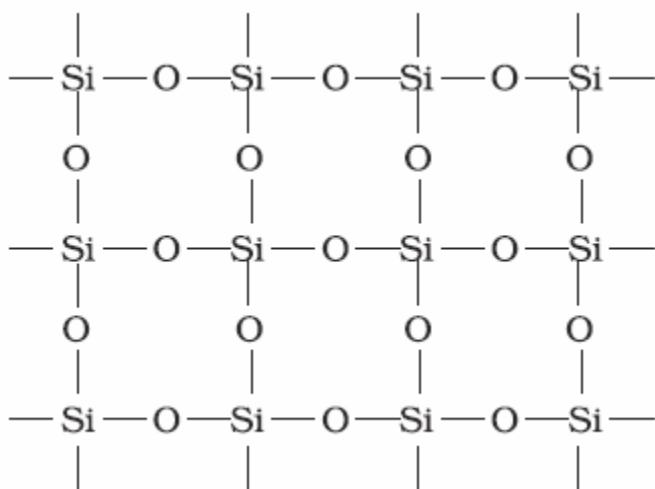
(c) Structures of carbon dioxide:

In CO₂ molecule carbon atom undergoes sp hybridisation. Two sp hybridised orbitals of carbon atom overlap with two p orbitals of oxygen atoms to make two sigma bonds while other two electrons of carbon atom are involved in pπ – pπ bonding with oxygen atom. This results in its linear shape [with both C–O bonds of equal length (115 pm)] with no dipole moment. The resonance structures are shown below:



(iii) Silicon dioxide (SO₂):

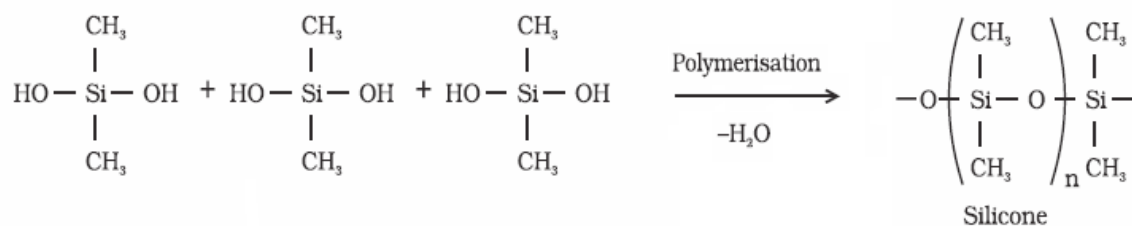
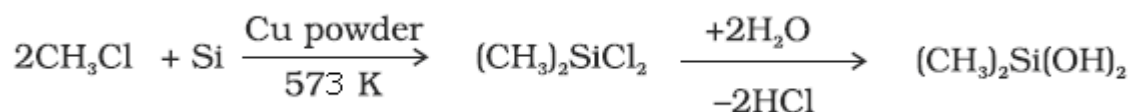
Silicon dioxide is a covalent, three dimensional network solid. In this structure each silicon atom is tetrahedrally surrounded by four oxygen atoms. Each oxygen atom in turn is covalently bonded to other silicon atoms. Each corner is shared with another tetrahedron. The entire crystal may be considered as giant molecule in which eight membered rings are formed with alternate silicon and oxygen atoms. Thus, there are no discrete SiO₂ units and it is a network solid.



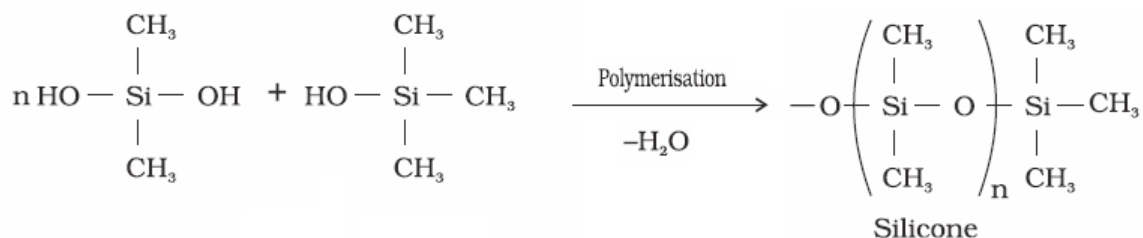
(iv) Silicones

Silicones are organosilicon polymeric compounds containing Si-O-Si linkages. These have the general formula $(R_2SiO)_n$ having (R_2SiO-) as a repeating unit. Starting materials for the manufacture of silicones are alkyl or aryl substituted silicon chlorides, $R_nSiCl_{(4-n)}$, where R is alkyl or aryl group.

When methyl chloride reacts with silicon in the presence of copper as a catalyst at a temperature 573K various types of methyl substituted chlorosilane of formula $MeSiCl_3$, Me_2SiCl_2 , Me_3SiCl with small amount of Me_4Si are formed. Hydrolysis of dimethyldichlorosilane, $(CH_3)_2SiCl_2$ followed by condensation polymerisation yields straight chain polymers.

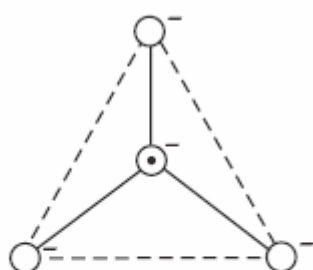


The chain length of the polymer can be controlled by adding $(CH_3)_3SiCl$ which blocks the ends.



(v) Silicates:

Basic building unit of all silicates is the tetrahedral SiO_4^{4-} ion. In silicates either the discrete unit is present or a number of such units are joined together via corners by sharing 1, 2, 3 or 4 oxygen atoms per silicate units. If all the four corners are shared with other tetrahedral units, three-dimensional network is formed.



- Silicon
- Oxygen

Representation of SiO_4^{4-} unit

(vi) Zeolites:

Aluminosilicates are formed when few Si atoms are replaced by Al atoms in three-dimensional network of silicon dioxide. Zeolites are widely used as a catalyst in petrochemical industries for cracking of hydrocarbons and isomerisation, e.g., ZSM-5 (A type of zeolite) used to convert alcohols directly into gasoline.

(7) Some important reasons:

(1) Atomic and ionic radii group 13 elements are smaller as compared to corresponding elements of group 2

Explanation: On moving from left to right in a periodic table, magnitude of nuclear charge increases but the electrons are added to the same shell. Since the electrons in the same shell do not screen each other, therefore, the electrons experience greater nuclear charge. In other words, effective nuclear charge increase and thus, size decrease. Therefore, the elements of group 13 have smaller size than the corresponding elements of second group.

(2) On moving down the group both atomic and ionic radii are expected to increase due to addition of new shells. But the atomic radius of Al is slightly more than that of Ga

Explanation: While going from Al ($Z = 13$) to Ga ($Z = 31$) there are ten elements of the first transition series of d-block from ($Z = 21$ to 30) which have electrons in the inner d-orbitals. The d-orbitals do not screen the nucleus effectively because of their shapes and poor penetration power. As a result, the effective nuclear charge in Ga becomes more than in Al and its atomic radius, therefore, decreases slightly.

(3) Ionization enthalpy decreases sharply from B to Al and then the ionization enthalpy of Ga is unexpectedly higher than that of Al

Explanation: The sharp decrease in ionization enthalpy from B to Al is due to increase in size. In case of Ga, there are ten d-electrons in its inner electronic configuration. Since, the d-electrons shield the nuclear charge less effectively than the s- and p-electrons, the outer electron is held strongly by the nucleus. As a result, the ionization enthalpy increases on moving from Al to Ga. Similar increase is also observed from In to Tl which is due to presence of 14 f-electrons in inner electronic configuration of Tl which have very poor shielding effect

(4) BCl_3 act as Lewis acid

Explanation: In trivalent state, the number of electrons around the central atom i.e. B in BCl_3 is only six. BCl_3 has two electrons less than the octet and therefore is electron deficient molecule. BCl_3 has a tendency to accept a pair of electrons to achieve stable electronic configuration and therefore, behave as Lewis acid

(5) AlCl_3 exist as dimer Al_2Cl_6 while BCl_3 exist as monomer, though both are electron deficient compounds

Explanation: In AlCl_3 , there are six electrons around Al atom and these are two less than the octet. In the dimeric structure, each aluminium atom completes its octet by accepting a lone pair of electron from the chlorine atom of another AlCl_3 molecule. Boron because of its small size cannot accommodate four large sized halide ions around it to form a dimer

(6) Anhydrous AlCl_3 covalent in nature but hydrated AlCl_3 is ionic

Explanation: In anhydrous AlCl_3 Al atom is linked to three Cl atoms by covalent bonds. This is due to the fact that a large amount of energy is needed to convert aluminium atom (Al) to aluminium ion (Al^{3+}). Therefore, it prefers to form covalent bonds with Cl atoms. However, when aluminium chloride is dissolved in water, it undergoes hydration as:



Hydration of Al_2Cl_6 is an exothermic process and a large amount of energy is released. Thus energy liberated during hydration process is responsible for the removal of three electrons from Al to form Al^{3+} ion. In fact, in dissolved

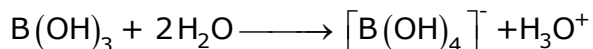
state, Al exists as $[\text{Al}(\text{H}_2\text{O}_6)]^{3+}$. Therefore, the hydrated aluminium chloride is ionic in nature.

(7) Basic character of hydroxides increase on moving down the group

Explanation: On moving down the group, the magnitude of the ionization energy decreases. As a result, the strength of the M–O bond also decreases accordingly. Therefore, its cleavage becomes easy resulting in the increased basic strength down the group

(8) Boric acid is not a protonic acid but a Lewis acid

Explanation: A species that accepts electron pair acts as Lewis acid. Boric acid accepts electrons from hydroxyl ion of water and a Lewis acid



(9) PbCl_4 is a good oxidizing agent

Explanation: An oxidizing agent reduces itself easily. Pb in PbCl_4 is in +4 oxidation state. Due to inert pair effect because Pb^{2+} is more stable than Pb^{4+} . Thus Pb^{4+} easily reduces to Pb^{2+} and thus PbCl_4 acts as good oxidizing agent

(10) Boron does not form B^{3+} ions

Explanation: Boron has a very small size. The sum of first three ionisation enthalpies ($\text{IE}_1 + \text{IE}_2 + \text{IE}_3$) is very high. Therefore, it cannot lose its three electrons to form B^{3+} ions.

(11) BCl_3 is more stable than TlCl_3

Explanation: B exhibits +3 oxidation state and can form stable BCl_3 . Thallium shows +3 oxidation state as well as +1 oxidation state but +1 oxidation state is more predominant stable than +3 oxidation state because of inert pair effect. Therefore, TlCl_3 is not as stable as BCl_3

(12) Molten aluminium bromide is a poor conductor of electricity

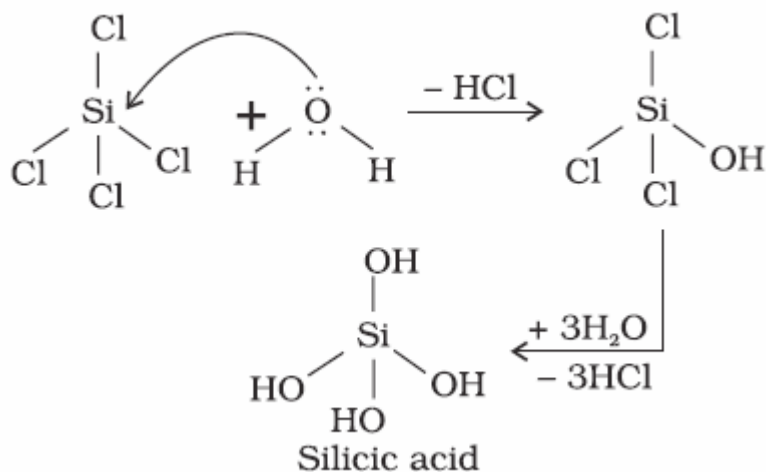
Explanation: AlBr_3 is predominantly a covalent compound. Even in molten state it does not have ions which can conduct electricity. Molten aluminium bromide is a poor conductor of electricity

(13) Tendency to catenation decreases in order $\text{C} \gg \text{Si} > \text{Ge} \approx \text{Sn} > \text{Pb}$

Explanation: Carbon has unique property of catenation i.e. property to form bonds with itself. Tendency to catenation depends upon the strength of M–M bond where $\text{M} = \text{C}, \text{Si}, \text{Ge}, \text{Sn}$ or Pb . If M–M bond formed is strong then catenation is favoured. As we go down the group the strength of M–M bond decreases due to increase in size of the atom and hence tendency to catenation decreases

(14) CCl_4 cannot be hydrolysed while SiCl_4 can be easily hydrolyzed

Explanation: SiCl_4 is readily hydrolysed by water because silicon has vacant 3d orbitals in its outermost shell. As a result, silicon can extend its coordination number beyond four. The vacant 3d orbitals of silicon accept electrons from water molecule and thus SiCl_4 can be easily hydrolysed



Carbon does not contain any d – orbitals in the valence shell. Carbon has only s and p – orbitals. As a result carbon cannot increase its coordination number beyond four and therefore, it cannot accept electrons from water molecules and eventually does not undergo hydrolysis

(15) Silanes are a few in number while alkanes are large in number

Explanation: Carbon has a maximum tendency for catenation due to strong C-C bonds. So, it forms a large number of alkanes. Silicon has a reduced tendency for catenation due to weak Si-Si bonds and hence it forms only a few silanes.

(16) Concentrated HNO_3 can be stored in an aluminium container but cannot be stored in a zinc container.

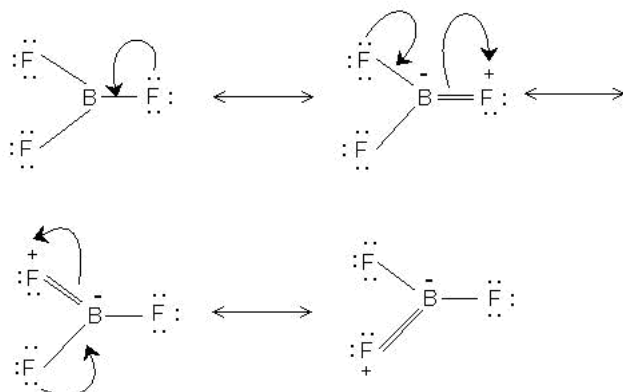
Explanation: In the presence of conc. HNO_3 , Al becomes passive, due to a thin protective layer of its oxide (Al_2O_3), which is formed on its surface and prevents further action between the metal and the acid. Therefore Al containers can be used for storing conc. HNO_3 . It cannot be stored in zinc vessels because zinc reacts with HNO_3 .

(17) BBr_3 is a stronger Lewis acid than BF_3

Explanation: BBr_3 is a stronger Lewis acid than BF_3 because of the back donation of electrons into empty 2p orbital of the B atom from filled p-orbital of the Br atom is much less than that by F atoms. This is due to larger size of the Br atom as compared to the F atom. As a result the electron deficiency of B is much higher in BBr_3 than that in BF_3 . Hence BBr_3 is a stronger Lewis acid than BF_3

(18) BF_3 exists whereas BH_3 does not exist

Explanation: BF_3 is trigonal planar molecule. Due to $p\pi - p\pi$ back bonding lone pair of electrons of F is back donated to B atom. This delocalization reduces the deficiency of electrons of boron thereby increasing the stability of BF_3 molecule. The mechanism is as follows:



Due to absence of lone pair of electrons on H atom this compensation does not occur in BH_3 .

In other words electron deficiency of B stays & hence it reduces its electron deficiency as BH_3 dimerises to form B_2H_6 .

(19) SiO_2 is solid but CO_2 is a gas at room temperature

Carbon is able to form $p\pi - p\pi$ bond with O atom and constitute a stable non-polar molecule $\text{O}=\text{C}=\text{O}$. Due to weak inter molecular force its boiling point is low and it is gas at room temperature.

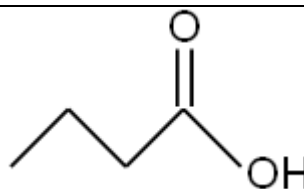
Si on the other hand is not able to form $p\pi - p\pi$ bond with O atoms because of its relatively large size. In order to complete its octet Si is linked to four O atoms around it by sigma bond & these constitute three dimensional network structures, which is responsible for its solid state.

Subject: Chemistry
Class: XI
Chapter: Organic Chemistry
Top concepts

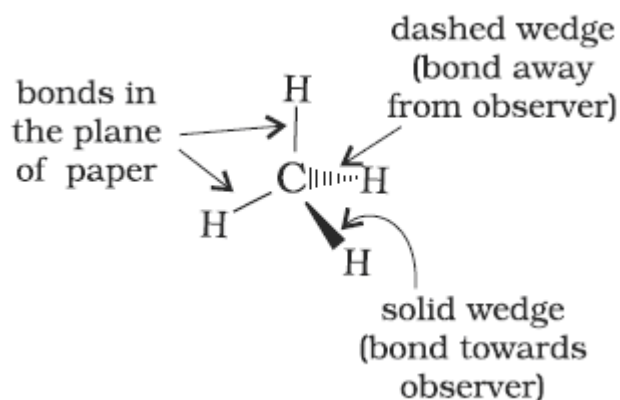
1. Study of chemistry of hydrocarbons and their derivatives is called organic chemistry
2. Reactivity of π -bonded compounds: The electron charge cloud of the π -bond is located above and below the plane of bonding atoms. These electrons are more exposed and therefore easily available to the electron seeking attacking reagents. Therefore π -bonds provide the most reactive centres in the molecules containing multiple bonds
3. Structural representation of Organic compounds:

	Example:
<p>Lewis structure :</p> <p>In this representation bond between atoms are represented by pairs of dots or lines and lone pairs on atoms are represented by a pairs of dots</p>	$ \begin{array}{cccc} & \text{H} & \text{H} & \text{H} & \ddot{\text{O}}: \\ & \vdots & \vdots & \vdots & \vdots \\ \text{H}: & \text{C} & : \text{C} & : \text{C} & : \text{C} & : \ddot{\text{O}} & : \text{H} \\ & \vdots & \vdots & \vdots & \vdots & \vdots & \\ & \text{H} & \text{H} & \text{H} & & & \end{array} $ <p>Or</p> $ \begin{array}{cccc} & \text{H} & \text{H} & \text{H} & \ddot{\text{O}}: \\ & & & & \\ \text{H} - & \text{C} & - \text{C} & - \text{C} & - \text{C} & - \ddot{\text{O}} - \text{H} \\ & & & & \vdots \\ & \text{H} & \text{H} & \text{H} & & \end{array} $
<p>Complete structural Formula:</p> <p>In this representation, a single bond is represented by single dash (-), double bond by double dash (=) and triple bond by triple dash (\equiv). Lone pair may or may not be shown on heteroatoms (for e.g. N, S, O etc.)</p>	$ \begin{array}{cccc} & \text{H} & \text{H} & \text{H} & \text{O} \\ & & & & \\ \text{H} - & \text{C} & - \text{C} & - \text{C} & - \text{C} & - \text{O} - \text{H} \\ & & & & \\ & \text{H} & \text{H} & \text{H} & \end{array} $
<p>Condensed structural formula:</p> <p>In this representation atoms are shown but bonds (dashes) between some or all atoms may be omitted and number of identical groups attached to an atom are indicated by a subscript</p>	$ \text{CH}_3\text{CH}_2\text{CH}_2\text{COOH} $ <p>or</p> $ \text{CH}_3(\text{CH}_2)_2\text{COOH} $
<p>Bond-line structural formula:</p>	

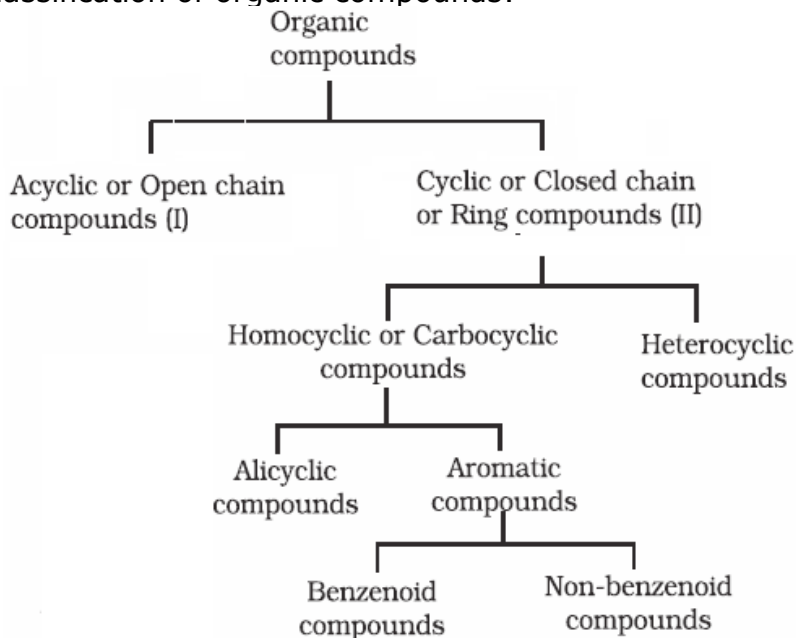
In this representation, carbon and hydrogen atoms are not shown and the lines representing carbon-carbon bonds are drawn in a zigzag fashion. Atoms other than carbon and hydrogen are written. The terminals denote methyl ($-\text{CH}_3$) groups (unless indicated otherwise by a functional group), while the line junctions denote carbon atoms bonded to appropriate number of hydrogens required to satisfy the valency of the carbon atoms



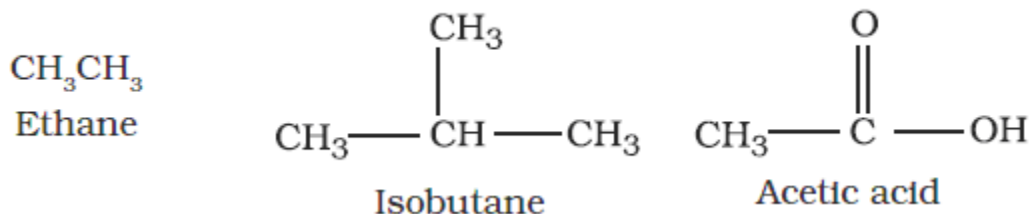
4. Three-dimensional representation of organic molecules: Solid-wedge is used to indicate a bond projecting out of the plane of paper, towards the observer. Dashed-wedge is used to depict the bond projecting out of the plane of the paper and away from the observer. Wedges are shown in such a way that the broad end of the wedge is towards the observer. The bonds lying in plane of the paper are depicted by using a normal line



5. Classification of organic compounds:



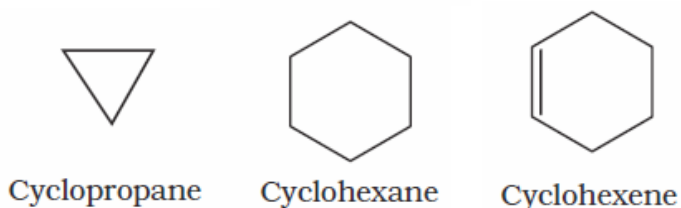
1. Acyclic or open chain compounds: Compounds contain open chain of carbon atoms in their molecule. Chains may be either branched chain or straight chain. Examples:



2. Closed chain or cyclic or ring compounds:


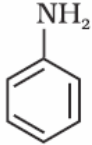
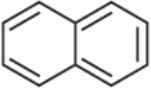

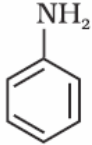
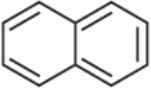


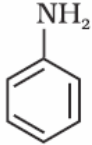
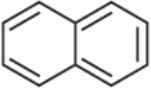
(a) Alicyclic or closed chain or ring compounds: These compounds contain ring of three or more carbon atoms in them. These compounds resemble aliphatic compounds in many of their properties.

Examples:



(b) Aromatic Compounds: These compounds have a cyclic system containing at least one benzene ring. Aromatic compounds have delocalised π - electron system

Classification of Aromatic compounds:

Benzenoid aromatic compounds	Non-benzenoid compound						
<p>These compounds contain benzene or other related ring compounds and exhibits aromaticity</p> <p>Examples:</p> <table style="width: 100%; text-align: center;"> <tr> <td></td> <td></td> <td></td> </tr> <tr> <td>Benzene</td> <td>Aniline</td> <td>Naphthalene</td> </tr> </table>				Benzene	Aniline	Naphthalene	<p>These compounds show aromaticity but do not contain any benzene ring</p> <p>Example:</p> <div style="text-align: center;">  <p>Tropolone</p> </div>
							
Benzene	Aniline	Naphthalene					

3. Heterocyclic compounds: In these compounds ring contains one or more atoms of N, O or S in addition to carbon atoms. Atom other than C (i.e. N, O or S) is called heteroatom

Examples:



Furan



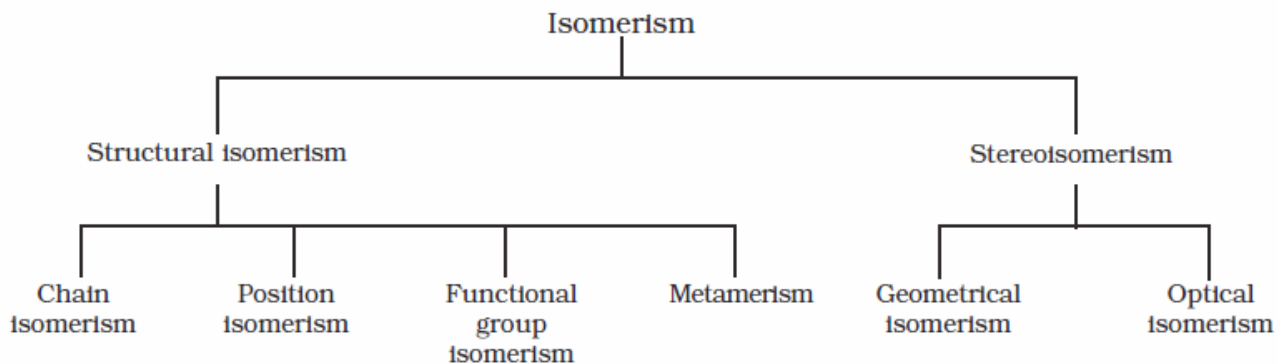
Thiophene

6. Functional group: An atom or group of atoms that determine the characteristic chemical properties of an organic compound

7. Homologous Series: A series of similarly constituted compounds in which members possess the same functional group and have similar chemical characteristics. Two consecutive members differ in their molecular formula by $-CH_2$ group. Different members of the series are known as homologous

8. Isomerism: The phenomenon of existence of two or more compounds possessing the same molecular formula but different properties is known as isomerism. Such compounds are called as isomers

Different types of isomerism:



(a) Structural isomerism: Compounds having same molecular formula, but different structures i.e. arrangement of atoms or groups of atoms within molecules are called structural isomers and phenomenon is called structural isomerism

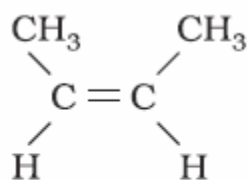
Classification of structural isomerism

1. Chain isomerism: Compounds having same molecular formula but different arrangement of carbon chains (skeletons), within the molecule are called chain isomers and the phenomenon is termed as chain isomerism

Example: Chain isomers of C_5H_{12} :

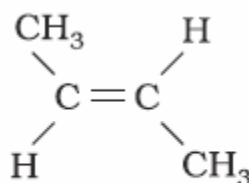
isomerism. Geometrical isomerism in alkenes is due to restricted rotation around carbon-carbon double bond.

Example: Geometrical isomers C_4H_8



cis-But-2-ene

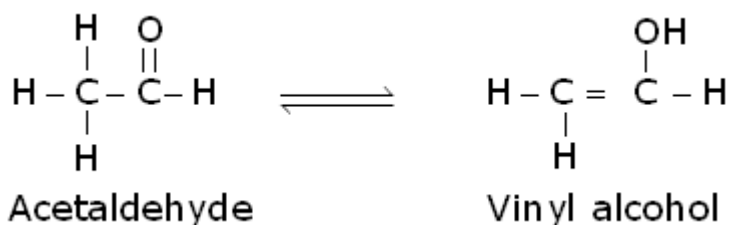
and



trans-But-2-ene

2. Optical isomerism: Stereoisomers involving the arrangement of substituents about an asymmetric carbon atom or atoms so that the various isomers differ in way they rotate a plane of polarized light are called optical isomers and the phenomenon is called optical isomerism

9. Tautomerism: It is a special type of functional isomerism in which the isomers differ in the arrangement of atoms but they exist in dynamic equilibrium with each other and this phenomenon is termed as tautomerism. Example: Acetaldehyde and vinyl alcohol are tautomers which exist in equilibrium as shown.

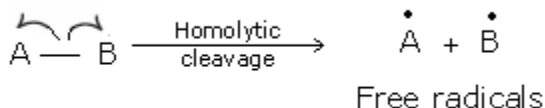


10. Substrate: In an organic reaction, reactant which supplies carbon to the new bond

11. Attacking reagent: In an organic reaction, chemical substance that attacks the organic molecule (substrate) and leads to the formation of product

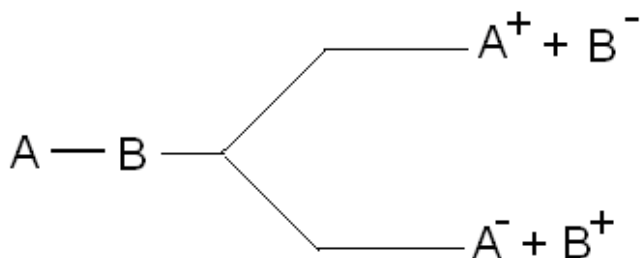
12. Reaction mechanism: Sequential account of each step, describing details of electron movement, energetics during bond cleavage and bond formation, and the rates of transformation of reactants into products (kinetics)

13. Homolytic cleavage of a covalent bond: Symmetrical cleavage of covalent bond between two atoms takes place resulting in the formation of neutral species (atoms or groups of atoms) having unpaired electrons called free radicals.



14. Heterolytic cleavage of a covalent bond: Unsymmetrical cleavage of covalent bond takes place resulting into two charged particles. The species that has a sextet at the carbon and is positively charged is called a carbocation. The species that has

a carbon atom with the shared pair of electrons and carrying a negative charge is called carbanion



Positively charged species is carbocation

Negatively charged species is carbanion

15. Reaction Intermediates: Species produced during cleavage of bonds

Name of the species	Order of stability of the species	Reason for stability order
Free radical: An atom or group of atoms having an unpaired electron.	$\cdot\text{CH}_3 < 1^\circ < 2^\circ < 3^\circ$	Order of stability of alkyl free radical can be explained on the basis of hyperconjugation. Larger the number of alkyl groups attached to the carbon atom carrying the odd electron, greater is the delocalization of the odd electron and hence more stable is the free radical. Thus, the tertiary free radical with three alkyl groups attached to the carbon atom carrying the odd electron is more stable than the secondary free radical containing two alkyl groups and so on.
Carbocation: Group of atoms which contain positively charged carbon having only six electrons (sextet of electrons)	$\text{CH}_3^+ < 1^\circ < 2^\circ < 3^\circ$	The alkyl group attached to positively charged carbon atom tends to release electron towards carbon due to +I inductive effect i.e., electron releasing. As a result, it decreases the +ve charge on the carbon atom but itself becomes somewhat positive. As a result, the positive charge on the carbon atom gets dispersed. The dispersal of charge results into stability. Therefore, more the number of alkyl groups, the greater will be the dispersal of charge and therefore, more stable will be the carbocation. Note: Triphenyl carbocation (Ph_3C^+) is most stable because of resonance stabilization among three phenyl groups
Carbanion:		The alkyl group attached to negatively charged carbon atom tends to release electron towards

Group of atoms which contain negatively charged carbon carrying negative charge	$\text{CH}_3^- > 1^\circ > 2^\circ > 3^\circ$	carbon due to +I inductive effect i.e., electron releasing. This increases electron density on the negatively charged carbon atom and hence makes it unstable. Larger the number of alkyl groups attached to the negatively charged carbon atom, greater will be the electron density on the carbon atom and lower will be its stability.
---	---	---

16. Different types of attacking reagent

Attacking Reagent	Examples
Free radical: An atom or group of atoms having an unpaired electron.	$\cdot\text{CH}_3, \cdot\text{C}_2\text{H}_5, \cdot\text{Cl}$ etc.
Electrophile: Positively charged or neutral species which are electron deficient and can accept a pair of electrons i.e. an electrophile takes away an electron pair	Positively charged: $\text{H}^+, \text{H}_3\text{O}^+, \text{CH}_3^+, \text{NO}_2^+$ etc.
	Neutral: $\text{AlCl}_3, \text{BF}_3, \text{SO}_3$ etc.
Nucleophile: Species that is electron rich and looks for electron deficient sites, i.e., nucleus loving or nucleus seeking (Nu) i.e. a nucleophile brings an electron pair with it	Negatively charged: $\text{X}^-, \text{OH}^-, \text{CN}^-, \text{RCOO}^-$ etc.
	Neutral: $\ddot{\text{N}}\text{H}_3, \text{R}_3\ddot{\text{N}}, \text{H}_2\ddot{\text{O}}, \ddot{\text{R}}\ddot{\text{O}}\text{H}, \ddot{\text{R}}\ddot{\text{O}}\text{R}$ etc.

17 Inductive Effect: The process of electron displacement along the chain of carbon atoms due to the presence of a polar covalent bond at one end of the chain is called inductive effect (or I-effect). It is a permanent effect.

NOTE: Inductive effect decreases on moving away from the atoms involved in the initial polar bond and becomes negligible from the fourth atom onwards

For comparing the relative effects, hydrogen is taken as standard and the atoms or groups can be classified into two categories:

(i) Atoms or groups of atoms having electron-attracting more than hydrogen are referred to as having -I (electron withdrawing or attracting) effect.

Example: $-\text{NO}_2 > -\text{CN} > -\text{COOH} > -\text{COOR} > -\text{F} > -\text{Cl} > -\text{Br} > -\text{I} > -\text{OH} > -\text{OCH}_3 > -\text{C}_6\text{H}_5 > \text{H}$

(ii) Atoms or groups of atoms having smaller electron attracting power than hydrogen are referred to as having +I (electron donating or repelling) effect.

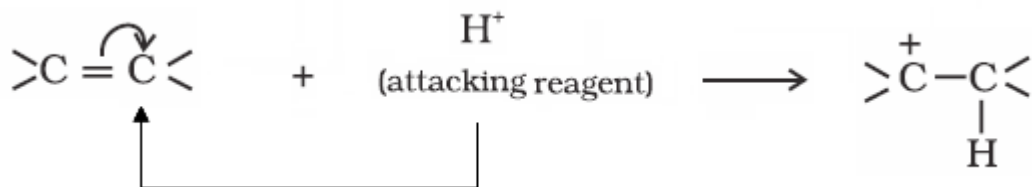
Example: $(\text{CH}_3)_3\text{C}- > (\text{CH}_3)_2\text{CH}- > \text{CH}_3\text{CH}_2- > \text{CH}_3-$

18. Electromeric Effect: It is a temporary effect which takes place between two atoms joined by a multiple bond, i.e., a double or a triple bond. This occurs at the requirements of the attacking reagent and involves instantaneous transfer of a shared pair of electrons of the multiple bonds to one of the linked atoms.

The electromeric effect is classified as +E effect and -E effect:

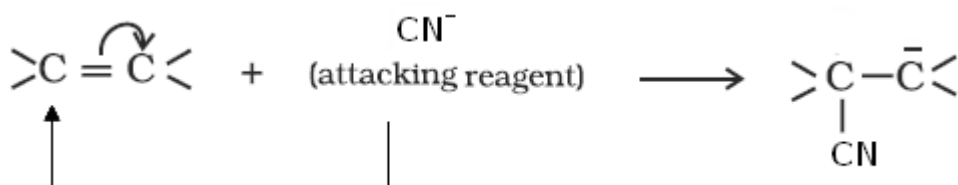
(i) When the π -electrons of the multiple bond are transferred to that atom to which the reagent gets attached, it is called +E (positive electromeric) effect.

Example:



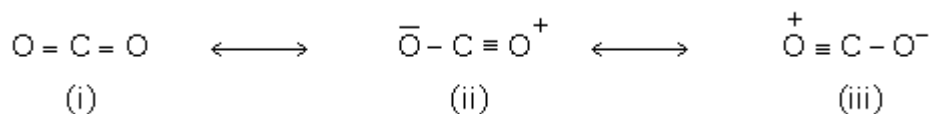
(ii) When the π -electrons of the multiple bond are transferred to that atom to which the attacking reagent does not get attached, it is called -E (negative electromeric) effect.

Example:



19. Resonance or Mesomeric Effect: If a molecule can be assigned two or more Lewis structures, none of which is capable of describing all the known properties of the compound, then the actual structure is intermediate or resonance hybrid of these structures. This phenomenon is called resonance. The various structures written are called resonating structures.

Example: Resonating structures of CO_2 are shown below



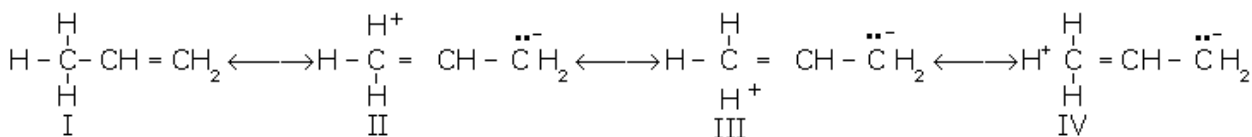
Resonance effect is classified as +R effect and -R effect:

(i) If a substituent has tendency to donate electrons to double bond or conjugated system, effect is called positive resonance effect or +R effect. Groups like -OH, -OR, -NH₂, -NHR, -NR₂, -Cl, -Br, etc show +R effect

(ii) If a substituent has tendency to withdraw electrons from a double bond or a conjugated system towards itself, effect is called negative resonance effect or -R effect. Groups like >C=O, -CHO, -CN, -NO₂, -COOR, etc show -R effect

20. Hyperconjugation: It involves delocalisation of (σ) electrons of C—H bond of an alkyl group directly attached to an atom of unsaturated system or to an atom with an unshared p orbital. The (σ) electrons of C—H bond of the alkyl group enter into partial conjugation with the attached unsaturated system or with the unshared p orbital. The interaction between the electrons of π systems (multiple bonds) and the adjacent σ bonds (single H-C bonds) of the substituent groups in organic compounds is called hyperconjugation. It is a permanent effect

Example: Hyperconjugation in propene



Since there is no bond between the α -carbon atom and one of the hydrogen atoms, the hyperconjugation is also called no-bond resonance. Although a free proton has been shown in the above structures, it is still bound quite firmly to the π -cloud and hence is not free to move.

Order of hyperconjugation: CH_3- > CH_3CH_2- > $(\text{CH}_3)_2\text{CH}-$ > $(\text{CH}_3)_3\text{C}-$

20. Filtration: Technique used separate an insoluble solid component of the mixture from the soluble component in a given solvent.

21. Recrystallisation: This method is based on the differences in the solubility of the organic compound and its impurities in a suitable solvent.

22. Simple Distillation: Method is used for the purification of liquids which boil without decomposition and contain non-volatile impurities. The simple distillation involves the heating of the liquid to its boiling point so that it is converted into vapours. On cooling the vapours, pure liquid is obtained and collected separately.

23. Fractional Distillation: Method is used to separate a mixture of two or more miscible liquids which have boiling points close to each other. Distillation is carried out by using fractionating columns. Fractionating column provides many surfaces for heat exchange between the ascending vapours and the descending condensed liquid.

24. Distillation under reduced pressure or Vacuum Distillation: Certain liquids have a tendency to decompose at a temperature below their boiling points. Such liquids cannot be purified by ordinary distillation. Under reduced pressure, the liquid will boil at a low temperature and the temperature of decomposition will not be reached.

25. Steam Distillation: Method is used for the separation and purification of liquid which is appreciably volatile in steam from non-volatile components of a mixture.

26. Differential Extraction: Method is used to separate a given organic compound present in aqueous solution by shaking with a suitable organic solvent in which the compound is more soluble than water. The basic requirement of the organic solvent is that it should be immiscible with water so that organic and water layers can be easily separated.

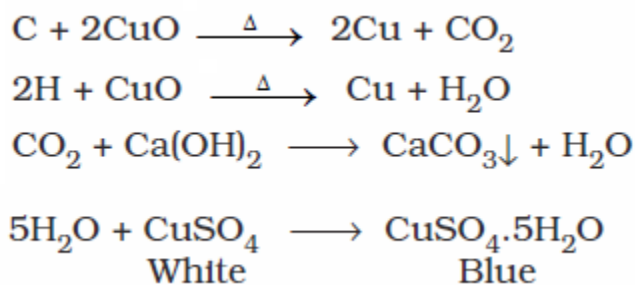
27. Chromatography: Technique of chromatography is based on the difference in the rates at which the components of a mixture move through a porous medium (called stationary phase) under the influence of some solvent or gas (called moving or mobile phase).

Based on the principle involved, chromatography is classified as:

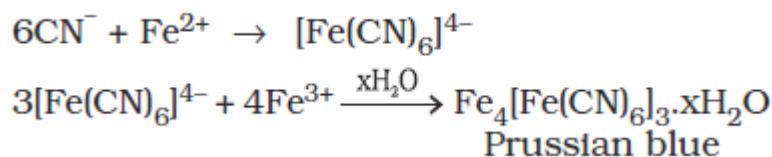
<p>(a) Adsorption Chromatography: It is based on the fact that different compounds are adsorbed on an adsorbent to different degrees. When a mobile phase is allowed to move over a stationary phase (adsorbent), the components of the mixture move by varying distances over the stationary phase because of the different adsorption tendencies.</p>	
<p>(i) Column Chromatography: In this method, a suitable adsorbent like alumina (Al_2O_3), silica (SiO_2) or Fuller's earth is packed as a column in a burette like long tube and this act as a stationary phase. The mixture to be separated is dissolved in a suitable solvent and the solution is poured on the top of the column of the adsorbent. The component which is adsorbed strongly gets adsorbed at the top. The other components with decreasing orders of their absorbability are held up at different zones down the column in the form of bands.</p>	<p>(ii) Thin Layer Chromatography: Method involves separation of mixture of substances over a thin layer of an adsorbent coated on glass plate. The solution of the mixture to be separated is applied as a small spot. After drying the glass plate is placed in a closed jar containing eluant. As the eluant rises up the plate, the components of the mixture move up along with the eluant to different distances depending on their degree of adsorption and separation takes place.</p>
<p>(b) Partition Chromatography: Method is based on differences in tendencies of substances to distribute or partition between stationary phase and mobile phase. Paper chromatography is a type of partition chromatography in which mixture to be separated is applied in form of a drop on the paper. This paper is then suspended in a suitable solvent or mixture of solvents. Solvent rises up by capillary action. Paper selectively retains different components according to their differing partition in two phases.</p>	

28. Qualitative Analysis of Organic Compounds: Analysis involving detection of all elements present in an organic compound

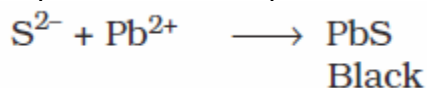
1. Detection of Carbon and Hydrogen: Carbon and hydrogen are detected by heating the compound with copper (II) oxide. Carbon present in the compound is oxidised to carbon dioxide and hydrogen to water. Carbon dioxide is tested with lime-water. Carbon dioxide makes lime water milky and develops turbidity. Water is tested with anhydrous copper sulphate, which becomes blue on absorbing moisture.



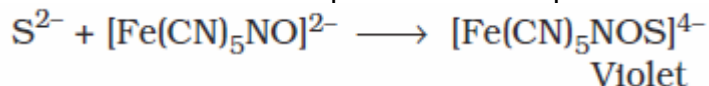
2. Detection of Nitrogen: The sodium fusion extract (prepared by boiling fused mass of compound and sodium metal in distill water) is boiled with FeSO₄ (iron (II) sulphate) and then acidified with concentrated sulphuric acid. The formation of Prussian blue colour confirms the presence of nitrogen.



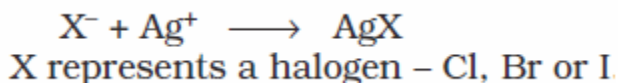
3. Detection of Sulphur: Two tests can be performed for detection of sulphur. (a) On adding lead acetate to acidified sodium fusion extract, formation of black precipitate confirms presence of sulphur



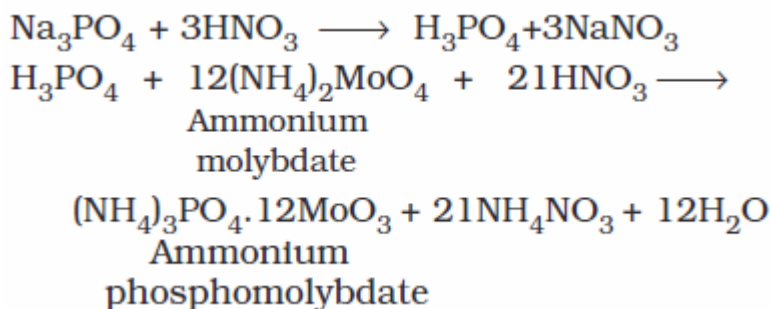
(b) On adding sodium nitroprusside to acidified sodium fusion extract, appearance of violet colour indicates presence of sulphur



4. Detection of Halogens: On adding silver nitrate solution to sodium fusion extract acidified with nitric acid white precipitate, soluble in ammonium hydroxide shows the presence of chlorine, a yellowish precipitate, sparingly soluble in ammonium hydroxide shows the presence of bromine and a yellow precipitate, insoluble in ammonium hydroxide shows the presence of iodine.



5. Detection of phosphorous: Compound containing phosphorous is heated with an oxidizing agent (sodium peroxide) which oxidises phosphorus to phosphate. The solution is boiled with nitric acid and then treated with ammonium molybdate. A yellow colouration or precipitate indicates the presence of phosphorus.



30. Quantitative Analysis of Organic Compounds: Analysis that involves determination of percentages of various elements present in a given compound

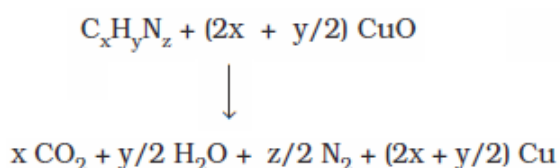
Principle involved in estimation of various elements:

1. Carbon and hydrogen: A known mass of the given dry organic compound is heated strongly with dry cupric oxide in an atmosphere of air or oxygen free from CO₂. The carbon and hydrogen of the organic compounds are oxidized to CO₂ and water vapour. Carbon dioxide produced is collected in potash bulbs (containing KOH solution) while vapour are absorbed in anhydrous calcium chloride tubes. From the amounts of carbon dioxide and water vapour produced, the percentage of carbon and hydrogen can be calculated.

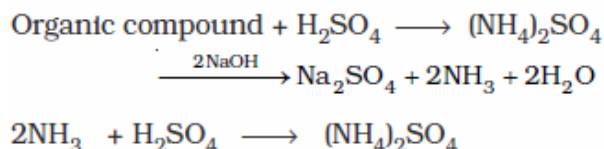


2. Nitrogen

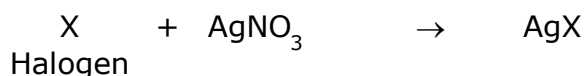
Duma's method: A known mass of an organic compound is heated with dry cupric oxide in an atmosphere of carbon dioxide. The carbon and hydrogen are oxidized to carbon dioxide and water respectively while nitrogen is set free. If any oxide of nitrogen is produced during this process, it is reduced to nitrogen by passing over heated copper gauze. The gaseous mixture is collected over an aqueous solution of KOH, when all the gases except nitrogen are absorbed. The volume of nitrogen produced is measured at room temperature and atmospheric pressure. From the volume of N₂ produced, percentage of nitrogen in the sample can be calculated.



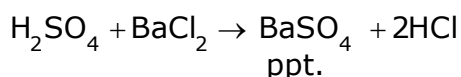
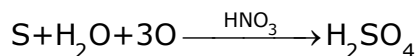
Kjeldahl's method: A known mass of the organic compound is heated with conc. H₂SO₄ so that nitrogen is quantitatively converted into ammonium sulphate. The resulting solution is then heated with excess of sodium hydroxide. The ammonia gas evolved is passed into a known but excess volume of standard acid (HCl or H₂SO₄). The acid left unused is estimated by titrating the solution with standard alkali. From the amount of acid left unused, the amount of acid used for neutralization of ammonia can be calculated.



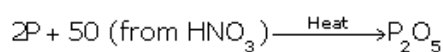
3. Halogens (Carius method): A known mass of the organic substance is heated with fuming HNO₃ in a Carius tube. The silver halide obtained is separated, washed, dried and weighed. From the weight of silver halide formed, the percentage of halogen can be calculated.



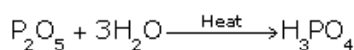
4. Sulphur (Carius method): A known mass of the organic compound is heated with sodium peroxide or fuming HNO₃ in a Carius tube when sulphur is quantitatively converted into sulphuric acid. It is then precipitated with barium chloride as barium sulphate. Precipitate is filtered, washed, dried and weighed. From the weight of BaSO₄ formed percentage of sulphur can be calculated.



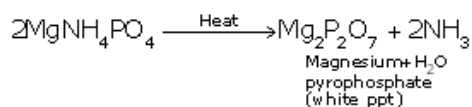
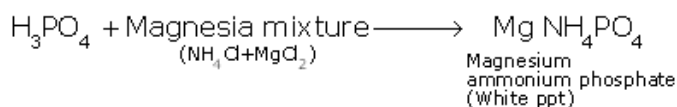
5. Phosphorus: Given organic compound is heated strongly with fuming nitric acid. The phosphorus in the compound is oxidized to phosphoric acid. It is treated with magnesia mixture (a solution containing magnesium chloride, ammonium chloride and a little of ammonia). A precipitate of magnesium ammonium phosphate $MgNH_4PO_4$ is formed. This is filtered, washed, dried and is then ignited to give magnesium pyrophosphate ($Mg_2P_2O_7$). From the weight of $Mg_2O_2O_7$, phosphorus can be estimated.



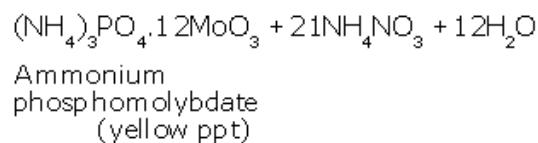
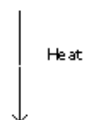
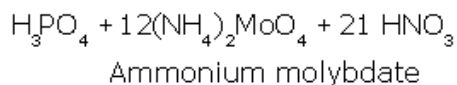
in organic
compound



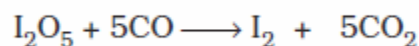
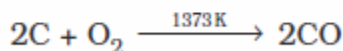
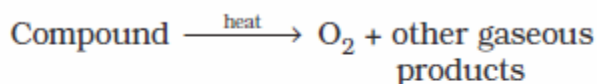
Phosphoric acid



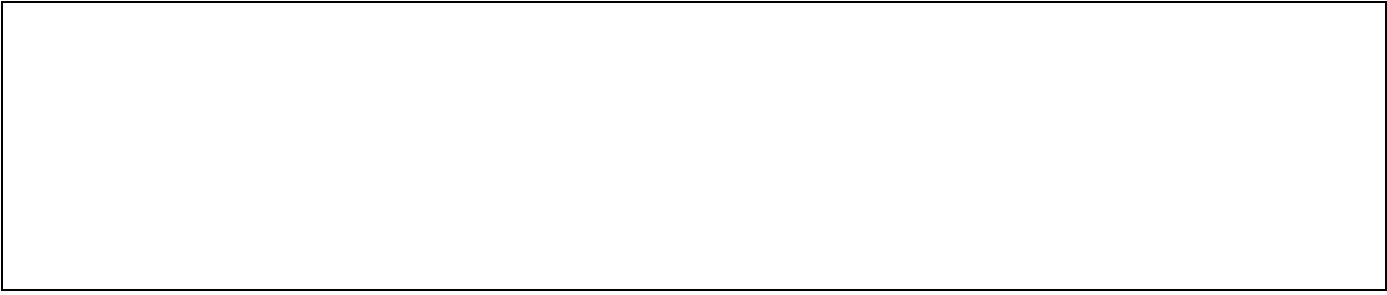
5. Phosphorus: A known mass of an organic compound is heated with fuming nitric acid. Due to this phosphorus present in the compound is oxidised to phosphoric acid. It is precipitated as ammonium phosphomolybdate, $(NH_4)_3PO_4 \cdot 12MoO_3$, by adding ammonia and ammonium molybdate. This is filtered, washed, dried and weighed. From the weight of ammonium phosphomolybdate, phosphorus can be estimated.



6. Oxygen: A definite amount of an organic compound is decomposed by heating in a stream of nitrogen gas. The mixture of gaseous products containing oxygen is passed over red hot coke when all the oxygen gets converted to carbon monoxide. The mixture is then passed through warm iodine pentoxide (I_2O_5) when carbon monoxide is oxidised to carbon dioxide producing iodine. The percentage of oxygen can be calculated from the amount of carbon dioxide or iodine produced.



Oxygen is generally estimated by subtracting the sum of the percentage of all other elements in the compound from 100.

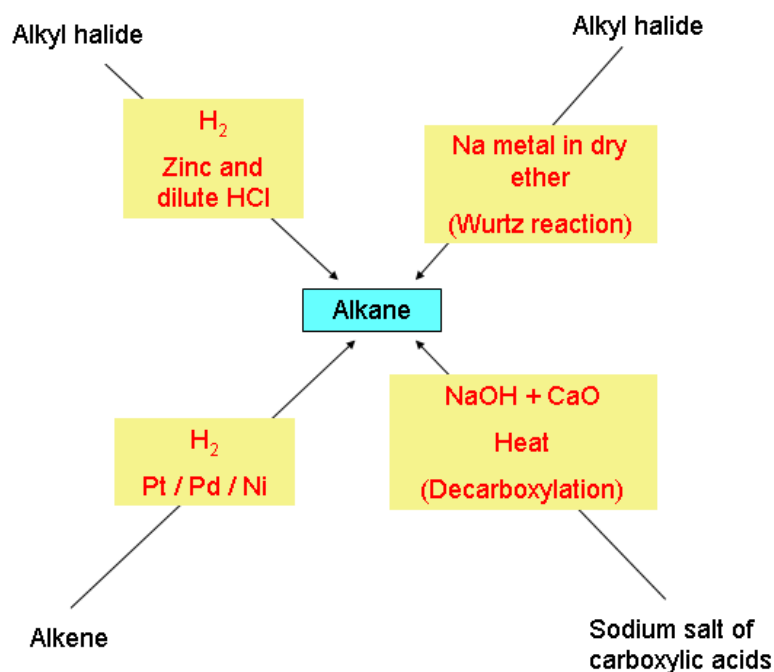


Class XI
Chemistry
Ch 13: Hydrocarbons

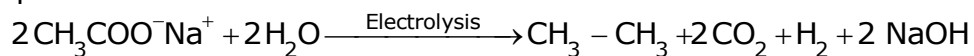
TOP Concepts:

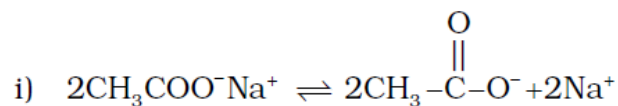
1. Alkanes: General formula: C_nH_{2n+2}

2. Preparation of alkanes:

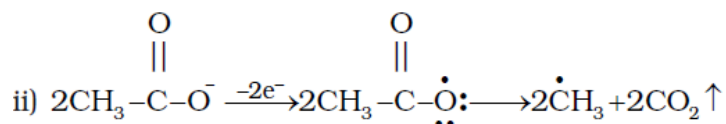


3. Kolbe's electrolytic method: Alkali metal salts of carboxylic acids undergo decarboxylation on electrolysis of their aqueous solutions and form hydrocarbons having twice the number of carbon atoms present in the alkyl group of the acid.





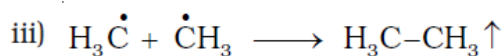
At anode:



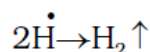
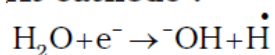
Acetate ion

Acetate
free radical

Methyl free
radical



iv) **At cathode :**



4. Physical properties of alkanes:

a. Physical state: C₁ to C₄ are gases, C₅ to C₁₇ are liquids and those containing 18 carbon atoms or more are solids at 298 K.

Reason: Hydrocarbons are non-polar because of the covalent nature of C-C and C-H bonds. Thus, they have weak van der Waals forces.

b. Boiling points: Alkanes have low boiling points because they have weak van der Waals forces.

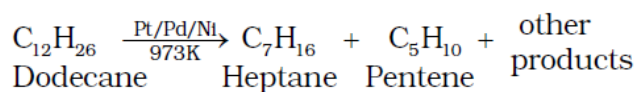
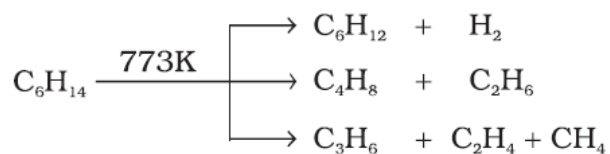
Higher alkanes with higher molecular masses have high boiling points than lower alkanes with low molecular masses. This is because the magnitude of van der Waals forces increases with increase in molecular mass or increase in surface area.

Boiling points of branched chain alkanes are lower than the corresponding straight chain alkanes. This is because with branching the molecules become more compact and hence the surface area decreases.

c. Melting points: The alkanes with even number of carbon atoms have higher melting points as compared to the immediately next lower alkane with odd number of carbon atoms. This is because alkanes with even number of carbon atoms have symmetrical structure and hence it results in closer packing in the crystal structure.

d. Solubility: Alkanes are soluble in non-polar solvents and are insoluble in polar solvents since alkanes are non-polar molecules.

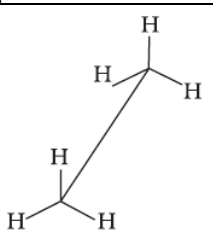
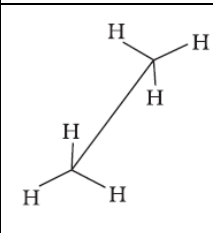
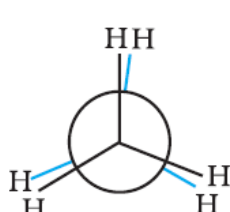
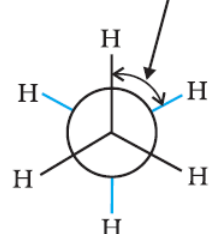
5. Chemical properties of alkanes:



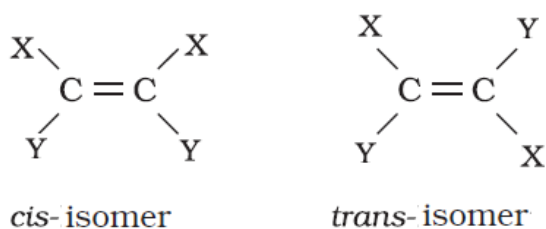
6. Conformations or conformers or rotamers or conformational isomers:

The spatial arrangements of atoms which can be converted into one another by rotation around a C-C single bond are called conformations or conformers or rotamers.

7. Conformations of ethane:

Conformations →	Eclipsed	Staggered	Skew
	Conformation in which hydrogen atoms attached to two carbons are as closed together as possible	Conformation in which hydrogen atoms are as far apart as possible	Any other intermediate conformation in between eclipsed and sawhorse
Sawhorse projection			
Newman projection		Angle of rotation or angle of torsion or dihedral angle 	

8. Geometrical isomers of alkenes:

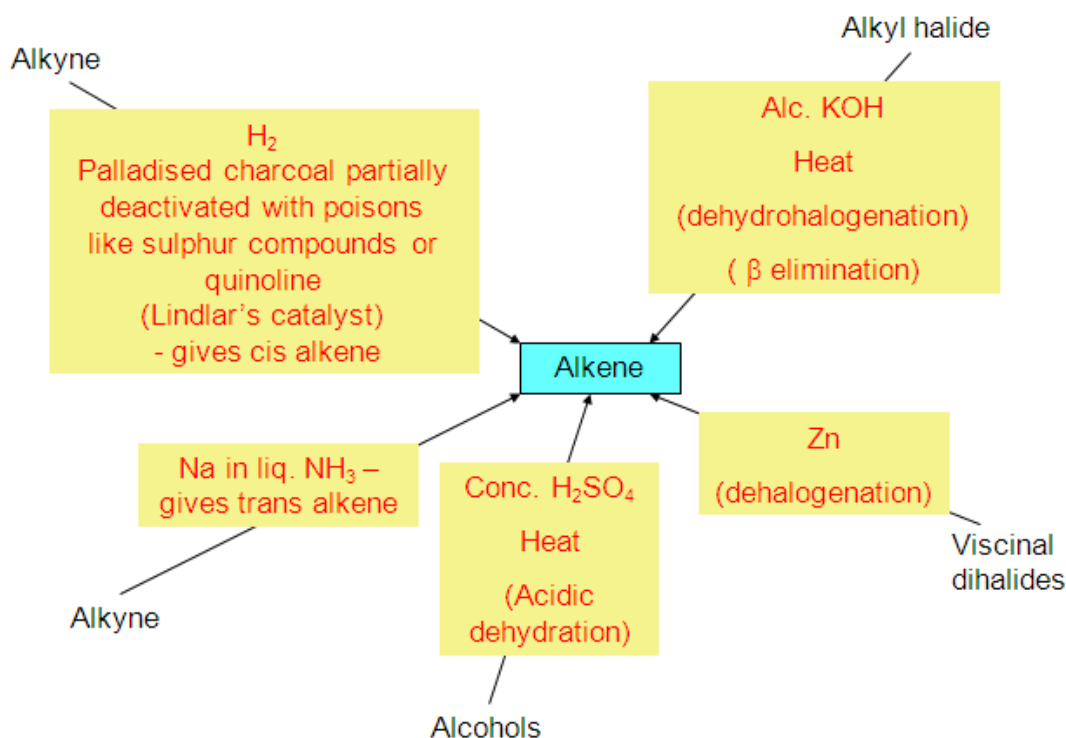


Trans isomer is non- polar and has zero dipole moment where as cis isomer is polar.

cis isomer has higher boiling points than trans isomer because it is polar.

In case of solids, cis isomer has a lower melting point than trans isomer. This is because trans isomer is symmetrical and fits well into the crystal lattice.

9. Preparation of alkenes:



10. Physical properties of alkenes:

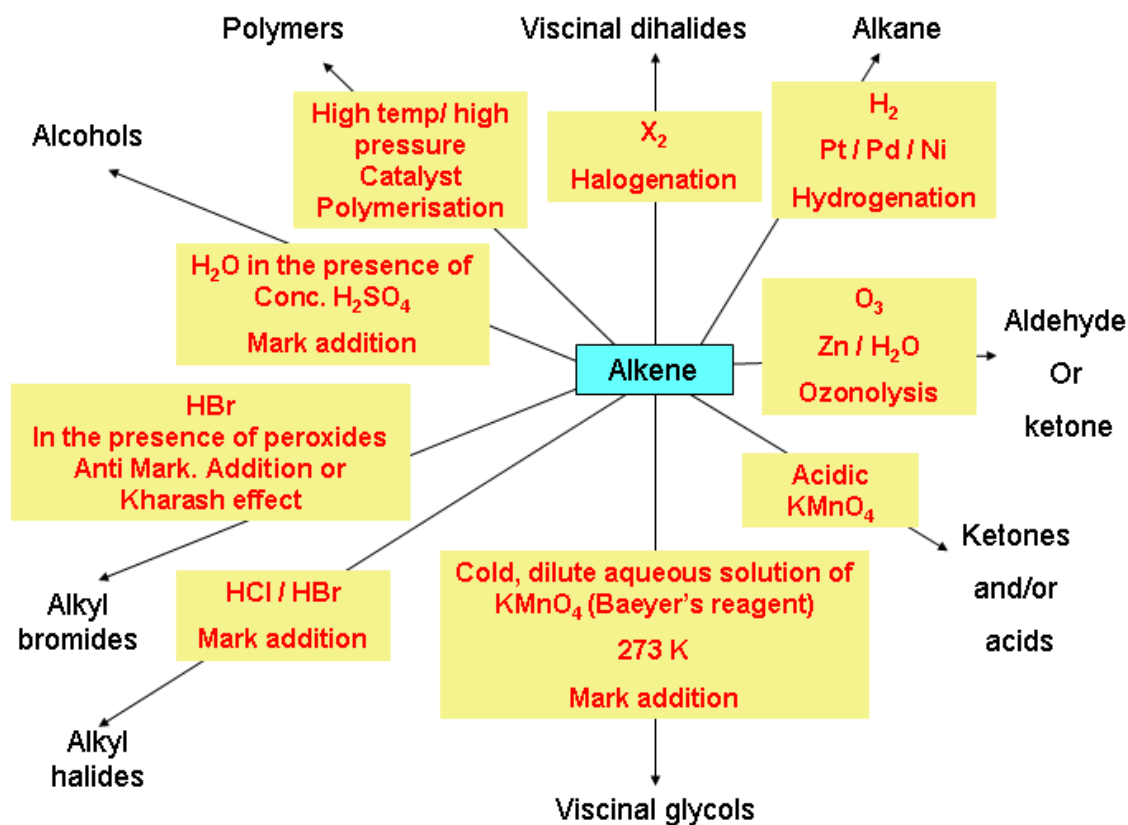
a. Physical state: The first three members are gases, the next fourteen are liquids and the higher ones are solids.

b. Boiling points: The boiling points of alkenes increase with the increase in molecular mass.

Boiling points of branched chain alkenes are lower than the corresponding straight chain alkenes. This is because with branching the molecules become more compact and hence the surface area decreases.

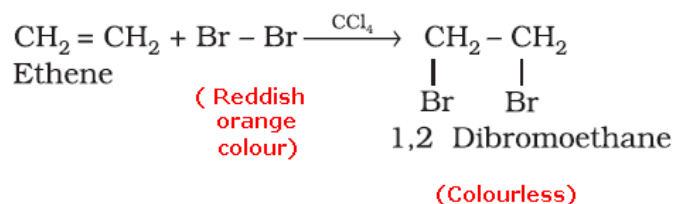
c. Solubility: Alkenes are soluble in non-polar solvents and are insoluble in polar solvents.

11. Chemical properties of alkenes:

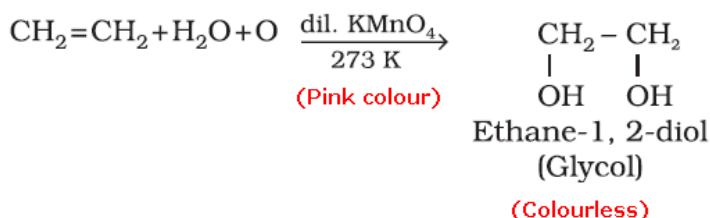


12. Test for Unsaturation: There are two tests to know whether a compound is unsaturated or not:

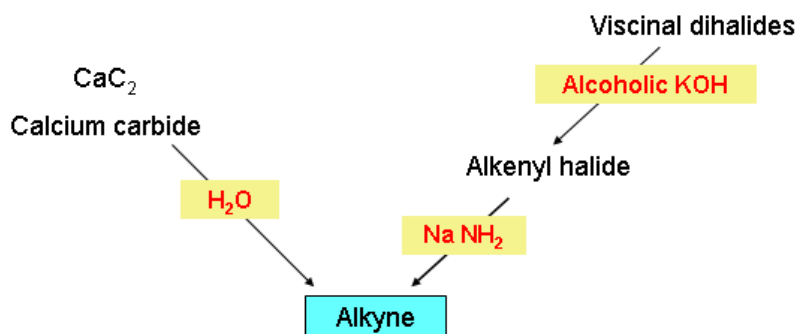
a. Bromine water test: When an unsaturated compound is reacted with bromine solution in carbon tetrachloride, the reddish orange colour of bromine solution in carbon tetrachloride is discharged.



b. Baeyer's reagent test: When an unsaturated compound is reacted with cold, dilute, aqueous solution of potassium permanganate (Baeyer's reagent, the pink colour of KMnO_4 solution is discharged.



13. Preparation of alkynes:



14. Physical properties of alkynes:

a. Solubility: Alkynes are weakly polar in nature. They are lighter than water and immiscible with water but soluble in organic solvents like ethers, carbon tetrachloride and benzene.

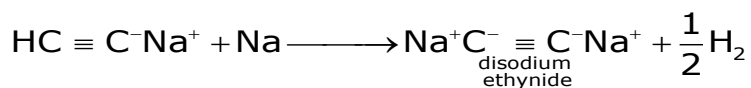
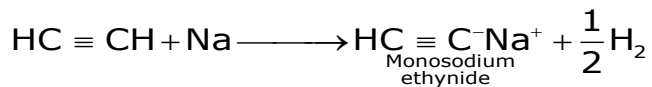
b. Melting and boiling point:

The melting and boiling points of the members of the alkyne family are slightly higher as compared to those of corresponding alkanes and alkenes. This is because alkynes have a linear structure and therefore their molecules are more closely packed as compared to alkanes and alkenes. The magnitude of attractive forces among them is higher and therefore, the melting and boiling points are also higher.

The melting point, boiling point and density of alkynes increase with increase in molar mass. This is because magnitude of van der Waals forces increases with increase in molecular mass or increase in surface area.

15. Acidic nature of alkynes: Hydrogen atoms of ethyne attached to triply bonded carbon atom are acidic in nature.

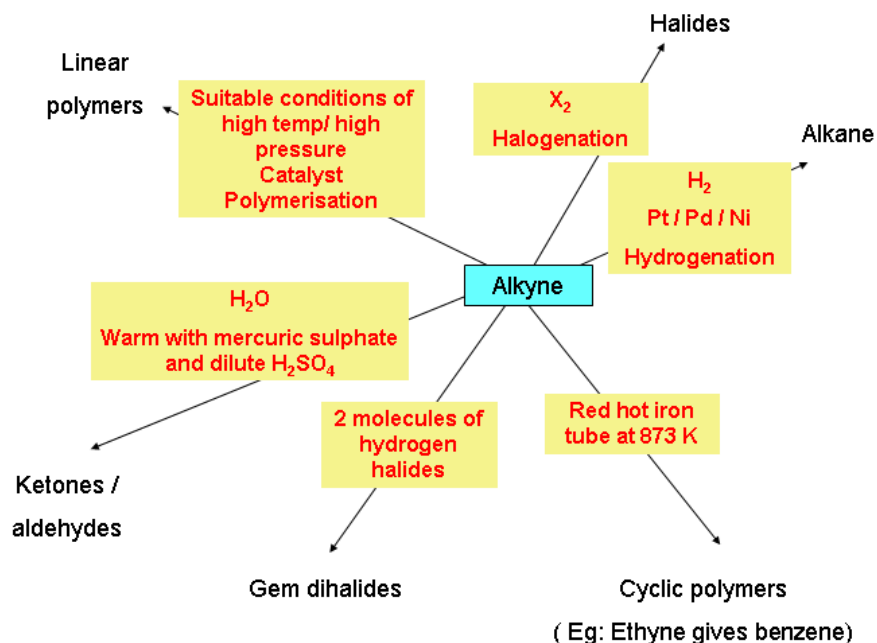
Example:



Reason:

	Alkyne	Alkene	Alkane
Hybridisation of carbon to which H atom is attached	sp	sp ²	sp ³
Percentage s character of carbon	50% (maximum)	66.6 %	25% (minimum)
Electronegativity of carbon atom	Highest	Less than alkyne and more than alkane	lowest
Extent of attraction of hydrogen atoms of C-H bonds towards C	Highest	Less than alkyne and more than alkane	lowest
Ease of liberation of H atoms as protons	Highest	Less than alkyne and more than alkane	lowest
Acidic character	Highest	Less than alkyne and more than alkane	lowest

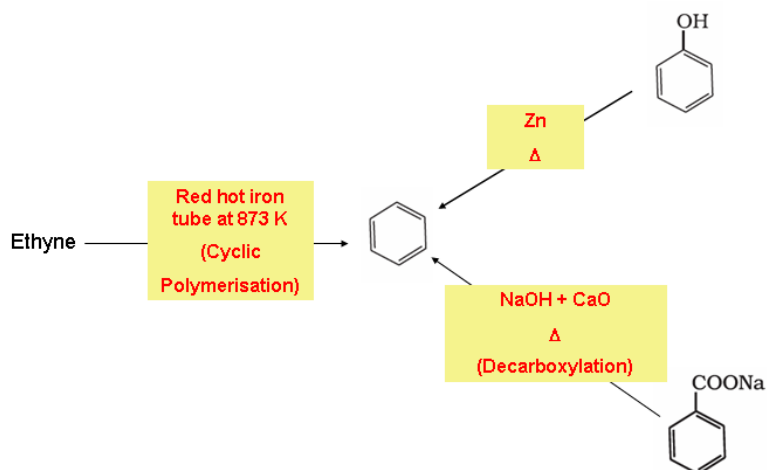
16. Chemical properties of alkynes:



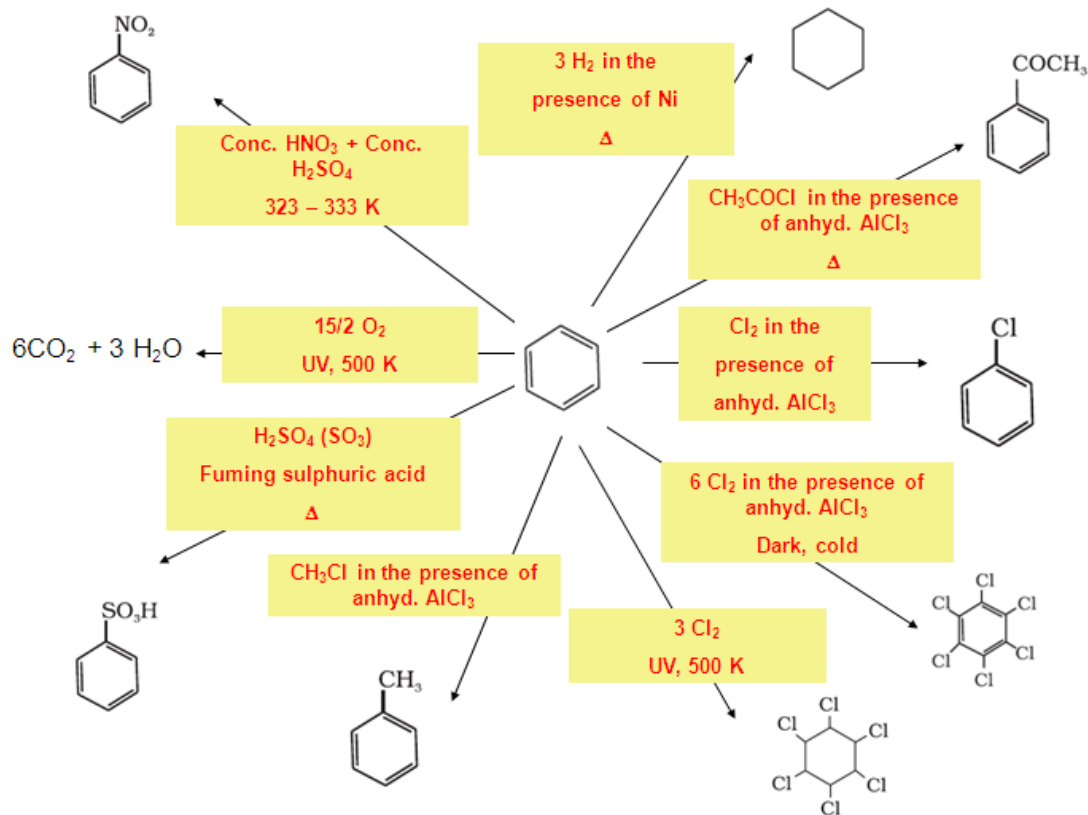
17. Conditions for aromaticity:

- Planarity: The ring should be planar. In simple words, all the carbons of the ring should be sp^2 hybridised.
- Complete delocalisation of the π electrons in the ring: The π electrons should be completely delocalized. In simple words, in the ring alternate single and double bonds should exist.
- Hückel Rule: Presence of $(4n + 2)$ π electrons in the ring where n is an integer ($n = 0, 1, 2, \dots$)

18. Preparation of benzene:



19. Chemical properties of benzene:



20. Directive influence of a functional group in monosubstituted benzene:

a. Ortho para directing groups: Groups which direct the incoming group to *ortho* and *para* positions.

Activating groups: Groups which increase the electron density on ortho and para positions of the benzene ring.

Example: $-\text{NH}_2$, $-\text{NHR}$, $-\text{NHCOCH}_3$, $-\text{OCH}_3$, $-\text{CH}_3$, $-\text{C}_2\text{H}_5$,

Deactivating groups: Groups which decrease the electron density on ortho and para positions of the benzene ring.

Example: $-\text{X}$

Halogens are moderately deactivating groups because they have strong $-I$ effect (electron withdrawing effect). Due to this the overall electron density on benzene ring decreases. It makes further substitution difficult. However, due to resonance the electron density on *o*- and *p*- positions is greater than that at the *m*-position. Hence, they are also *o*- and *p*- directing groups.

b. Meta directing groups: The groups which direct the incoming group to *meta* position.

Example: $-\text{NO}_2$, $-\text{CN}$, $-\text{CHO}$, $-\text{COR}$, $-\text{COOH}$, $-\text{COOR}$, $-\text{SO}_3\text{H}$

Class XI
Chemistry
Ch14: Environmental Chemistry

TOP Concepts:

1. **Environmental pollution:** It is the effect of undesirable changes in our surroundings that have harmful effects on plants, animals and human beings.
2. **Pollutant:** A substance, which causes pollution, is known as pollutant.
3. Pollutants can be solid, liquid or gaseous substances present in greater concentration than in natural abundance.
4. Pollutants can be natural or anthropogenic:
 - a. **Natural pollutants:** These are produced due to natural happenings like volcano eruptions etc.
 - b. **Anthropogenic pollutants:** These are produced due to human activities.
5. Pollutants can be biodegradable or non – biodegradable:
 - a. **Biodegradable pollutants:** These are the pollutants which rapidly break down by natural processes. Example: discarded vegetables
 - b. **Non – biodegradable pollutants:** These are the pollutants which are slowly degradable, and remain in the environment in an unchanged form for many decades. For example: DDT, plastic materials, heavy metals, many chemicals, nuclear wastes etc
6. Environmental pollution is of three types:
 - a. Atmospheric pollution
 - i. Tropospheric pollution
 - ii. Stratospheric pollution
 - b. Water pollution
 - c. Soil and land pollution
7. Atmospheric pollution occurs when the concentration of a normal component of the air or a new chemical substance added or formed in air builds up to undesirable proportions causing harm to humans, other animals, vegetation and materials.
8. **Troposphere:** The lowest region of atmosphere in which the human beings along with other organisms live is called troposphere. It extends up to the height of ~ 10 km from sea level.
9. **Stratosphere:** Above the troposphere, between 10 and 50 km above sea level lies stratosphere.
10. **Tropospheric pollution:** Is because of two types of pollutants:

a. **Gaseous air pollutants:** These are oxides of sulphur, nitrogen and carbon, hydrogen sulphide, hydrocarbons, ozone and other oxidants.

b. **Particulate pollutants:** Particulate pollutants are the minute solid particles or liquid droplets in air. These are present in vehicle emissions, smoke particles from fires, dust particles and ash from industries. Examples of particulate pollutants are dust, mist, fumes, smoke, smog etc.

11. Oxides of sulphur as pollutant:

Sources: Burning of fossil fuels containing sulphur

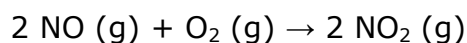
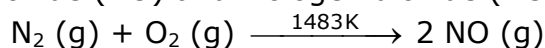
Harmful effects:

- Causes respiratory diseases e.g., asthma, bronchitis, emphysema in human beings.
- Sulphur dioxide causes irritation to the eyes, resulting in tears and redness.
- High concentration of sulphur dioxide leads to stiffness of flower buds which eventually fall off from plants.

12. Oxides of nitrogen as pollutant:

Sources:

- At high altitudes when lightning strikes, dinitrogen and dioxygen combine to form oxides of nitrogen.
- Burning of fossil fuel in an automobile engine, at high temperature, dinitrogen and dioxygen combine to yield significant quantities of nitric oxide (NO) and nitrogen dioxide (NO₂).



Harmful effects:

- Damage the leaves of plants and retard the rate of photosynthesis
- Nitrogen dioxide is a lung irritant that can lead to an acute respiratory disease in children
- It is toxic to living tissues also
- Nitrogen dioxide is also harmful to various textile fibres and metals

13. Hydrocarbons as pollutant:

Source: Incomplete combustion of fuel used in automobiles

Harmful effects:

- Hydrocarbons are carcinogenic, i.e., they cause cancer
- They harm plants by causing ageing, breakdown of tissues and shedding of leaves, flowers and twigs

14. Oxides of carbon as pollutant:

a. Carbon monoxide:

Source:

- Incomplete combustion of carbon of coal, firewood, petrol, etc
- By automobile exhaust

Harmful effects:

It is highly poisonous to living beings because of its ability to block the delivery of oxygen to the organs and tissues. It binds to haemoglobin to form carboxyhaemoglobin, which is about 300 times more stable than the oxygen-haemoglobin complex. In blood, when the concentration of carboxyhaemoglobin reaches about 3–4 per cent, the oxygen carrying capacity of blood is greatly reduced. This oxygen deficiency, results into headache, weak eyesight, nervousness and cardiovascular disorder.

b. Carbon dioxide:

Source:

- Respiration
- Burning of fossil fuels for energy
- By decomposition of limestone during the manufacture of cement
- By volcanic eruptions
- Deforestation

Harmful effects:

- Causes global warming

15. **Green house effect:** About 75 % of the solar energy reaching the earth is absorbed by the earth's surface, which increases its temperature. The rest of the heat radiates back to the atmosphere. Some of the heat is trapped by gases such as carbon dioxide, methane, ozone, chlorofluorocarbon compounds (CFCs) and water vapour in the atmosphere. Thus, they add to the heating of the atmosphere. This causes global warming.

This trapping of the sun's heat near the earth's surface and keeping it warm is called natural greenhouse effect. It maintains the temperature and makes the earth perfect for life.

If the amount of carbon dioxide crosses the delicate proportion of 0.03 per cent, the natural greenhouse balance may get disturbed. This may lead to global warming.

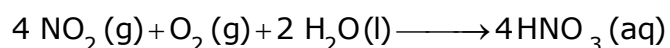
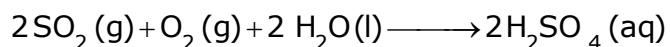
16. **Green house:** In a greenhouse, visible light passes through the transparent glass and heats up the soil and the plants. The warm soil and plants emit infrared radiations. Since glass is opaque to infrared (heat) radiations, it partly reflects and partly absorbs these radiations. This mechanism keeps the energy of the sun trapped in the greenhouse.

17. **Global warming:** An increase in the average temperature of the earth's atmosphere (especially a sustained increase that causes climatic changes) which may be caused by additional heat being trapped by greenhouse gases.

18. **Acid rain:** Normally rain water has a pH of 5.6 due to the presence of H⁺ ions formed by the reaction of rain water with carbon dioxide present in the atmosphere.



Source: Burning of fossil fuels (which contain sulphur and nitrogenous matter) such as coal and oil in power stations and furnaces or petrol and diesel in motor engines produce sulphur dioxide and nitrogen oxides. SO_2 and NO_2 after oxidation and reaction with water are major contributors to acid rain, because polluted air usually contains particulate matter that catalyses the oxidation.



Harmful effects:

- Harmful for agriculture, trees and plants as it dissolves and washes away nutrients needed for their growth.
- Causes respiratory ailments in human beings and animals.
- Affects plant and animal life in aquatic ecosystem when acid rain falls and flows as ground water to reach rivers, lakes etc.
- Corrodes water pipes resulting in the leaching of heavy metals such as iron, lead and copper into the drinking water.
- Damages buildings and other structures made of stone or metal. The Taj Mahal in India has been affected by acid rain.

19. **Particulates in the atmosphere may be viable or non-viable:**

- a. Viable are minute living organisms that are dispersed in the atmosphere. Example: bacteria, fungi, moulds, algae etc.
- b. Non-viable particulates may be classified as:
 - i. Smoke particulates: consist of solid or mixture of solid and liquid particles formed during combustion of organic matter. Example: cigarette smoke, smoke from burning of fossil fuel, garbage and dry leaves, oil smoke etc.
 - ii. Dust: composed of fine solid particles (over $1\mu\text{m}$ in diameter), produced during crushing, grinding and attribution of solid materials. Sand from sand blasting, saw dust from wood works, pulverized coal, cement and fly ash from factories, dust storms etc., are some typical examples of this type of particulate emission.
 - iii. Mists: Are produced by particles of spray liquids and by condensation of vapours in air. Example: sulphuric acid mist and herbicides and insecticides that miss their targets and travel through air and form mists.
 - iv. Fumes: Are generally obtained by the condensation of vapours during sublimation, distillation, boiling and several other chemical reactions. Generally, organic solvents, metals and metallic oxides form fume particles.

20. **Smog:** Smoke is a mixture of smoke, dust particles and small drops of fog.

21. **Smog is of two types:**

Classical Smog	Photochemical Smog
1. It occurs in cool humid climate.	1. It occurs in warm, dry and sunny climate.
2. It is a mixture of smoke, fog & sulphur dioxide.	2. Components of photochemical smog result from the action of sunlight on unsaturated hydrocarbons & oxides of nitrogen produced by automobiles & factories.
3. It is also called reducing smog.	3. It is also called oxidizing smog.

22. **Formation of photochemical smog:**

Burning of fossil fuels



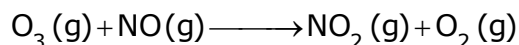
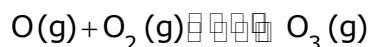
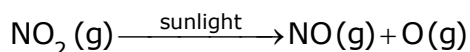
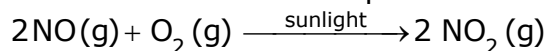
Emission of a variety of pollutants (hydrocarbons and nitric oxide to troposphere



At high levels, leads to



Chain reaction between pollutants and sunlight:



NO_2 and O_3 are strong oxidising agents and can react with the unburnt hydrocarbons in the polluted air to produce chemicals such as formaldehyde, acrolein and peroxyacetyl nitrate (PAN).

23. **Effects of photochemical smog:**

- Ozone and PAN act as powerful eye irritants.
- Ozone and nitric oxide irritate the nose and throat and their high concentration causes headache, chest pain, and dryness of the throat, cough and difficulty in breathing.

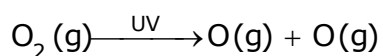
- Photochemical smog leads to cracking of rubber and extensive damage to plant life.
- It also causes corrosion of metals, stones, building materials, rubber and painted surfaces.

24. Control of photochemical smog:

- Use of catalytic converters in automobiles, which prevent the release of nitrogen oxide and hydrocarbons to the atmosphere.
- Certain plants e.g., Pinus, Juniparus, Quercus, Pyrus and Vitis can metabolise nitrogen oxide and therefore, their plantation could help in this matter.

25. Stratospheric pollution is basically due to ozone layer depletion.

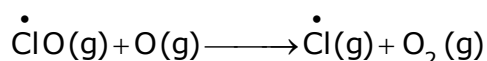
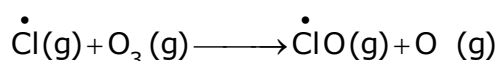
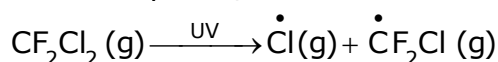
26. Formation of ozone in stratosphere:



27. Depletion of ozone layer:

Release of chlorofluorocarbon compounds (CFCs), also known as freons lead to their mixing with the normal atmospheric gases and eventually reach the stratosphere.

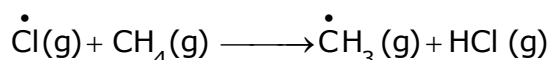
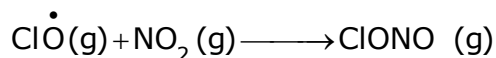
In stratosphere,



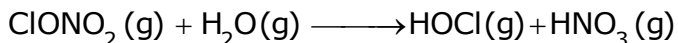
This way, the chlorine radicals are continuously regenerated and cause the breakdown of ozone layer.

28. Ozone hole over Antarctica:

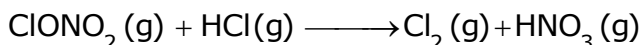
In summer season, nitrogen dioxide and methane react with chlorine monoxide and chlorine atoms forming chlorine sinks, preventing much ozone depletion.



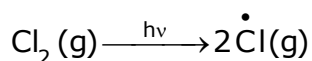
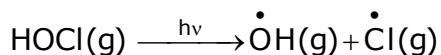
In winter, special type of clouds called polar stratospheric clouds are formed over Antarctica. These polar stratospheric clouds provide surface on which chlorine nitrate formed gets hydrolysed to form hypochlorous acid.



It also reacts with hydrogen chloride produced to give molecular chlorine.



When sunlight returns to the Antarctica in the spring, the sun's warmth breaks up the clouds and HOCl and Cl₂ are photolysed by sunlight.



The chlorine radicals thus formed, initiate the chain reaction for ozone depletion.

29. Effects of depletion of the ozone layer:

With the depletion of ozone layer, more UV radiation filters into troposphere. UV radiations lead to:

- Ageing of skin, cataract, sunburn and skin cancer etc in human beings
- Killing of many phytoplanktons
- Damage to fish productivity
- Affect the plant proteins which lead to the harmful mutation of cells
- Increases the evaporation of surface water through the stomata of the leaves and decreases the moisture content of the soil
- Increase in UV radiations damage paints and fibres, causing them to fade faster

30. Water pollution:

Major water pollutants	Sources	Harmful effects
Pathogens (Micro-organisms)	Domestic sewage	Cause gastrointestinal diseases.
Organic wastes (leaves, grass, trash)	Domestic sewage, animal excreta and waste, decaying animals and plants, discharge from food processing factories	Lead to decrease in concentration of dissolved oxygen in water and lead to death of aquatic life
Plant nutrients	Chemical fertilizers	
Toxic heavy metals (cadmium, mercury, nickel)	Industries and chemical factories	Can damage kidneys, central nervous system, liver etc
Sediments	Erosion of soil by agriculture and strip mining	
Pesticides	Chemicals used for	Lead to eutrophication

(insecticides, herbicides, fungicides)	killing insects, fungi and weeds	
Radioactive substances	Mining of uranium containing minerals	
Heat	Water used for cooling in industries	

31. **Biochemical Oxygen Demand (BOD):** The amount of oxygen required by bacteria to break down the organic matter present in a certain volume of a sample of water is called BOD.

32. **Eutrophication:** The process in which nutrient enriched water bodies support a dense plant population, which kills animal life by depriving it of oxygen and results in subsequent loss of biodiversity, is known as eutrophication.

33. **Some constituents of drinking water:**

Constituent	Maximum concentration	Harmful effects of higher concentration
Fluoride	1 ppm or 1 mg dm ⁻³	Causes brown mottling of teeth
Lead	50 ppb	Can damage kidney, liver, reproductive system etc
Sulphate	500 ppm	Causes laxative effect
Nitrate	50 ppm	Causes disease such as methemoglobinemia ('blue baby' syndrome)
Metals		
Fe	0.2 ppm	
Al	0.05 ppm	
Mn	0.2 ppm	
Cu	3.0 ppm	
Zn	5.0 ppm	
Cd	0.005 ppm	

34. **Pesticides:** They are basically synthetic toxic chemicals with ecological repercussions.

35. **Herbicides:** They are used to kill weeds or undesirable vegetation.
Example: sodium chlorate (NaClO_3), sodium arsenite (Na_3AsO_3)

36. **Strategies to control environmental pollution:**

a. Water management

- Segregate the water as biodegradable and non- biodegradable waste:
 - Biodegradable waste:
 - ✓ Generated by cotton mills, food processing units, paper mills, and textile factories.
 - ✓ Management: are deposited in landfills and are converted into compost
 - Non – biodegradable water:
 - ✓ Generated by thermal power plants which produce fly ash; integrated iron and steel plants which produce blast furnace slag and steel melting slag
 - ✓ Management:
 - Recycling
 - Toxic wastes are usually destroyed by controlled incineration

c. Green chemistry: Green chemistry is a strategy to design chemical processes and products that reduces or eliminates the use and generation of hazardous substances. The chemical reactions should be such that the reactants are fully converted into useful environmental friendly products by using an environment friendly medium so that there would be no chemical pollutants introduced in the environment.

37. **Green chemistry in daily life:**

Purpose	Earlier	Now
Dry cleaning of clothes	Tetrachloroethene ($\text{Cl}_2\text{C}=\text{CCl}_2$) which contaminates the ground water	Liquefied carbon dioxide, with a suitable detergent
Bleaching of paper	Chlorine gas	Hydrogen peroxide (H_2O_2) with suitable catalyst