Chapter 04

Liquids and Solids

Intermolecular Forces

Definition

The force of attraction between the molecules is called intermolecular force. For example, the force of attraction between HCl molecules.

Types of Intermolecular Forces

Following are the four types:

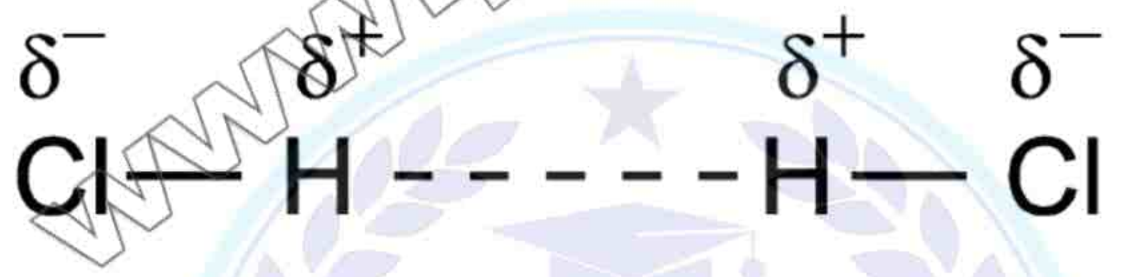
1. Dipole-dipole forces
2. Ion-dipole forces
3. Dipole-induced dipole forces
4. Instantaneous dipole-induced dipole forces or London dispersion forces

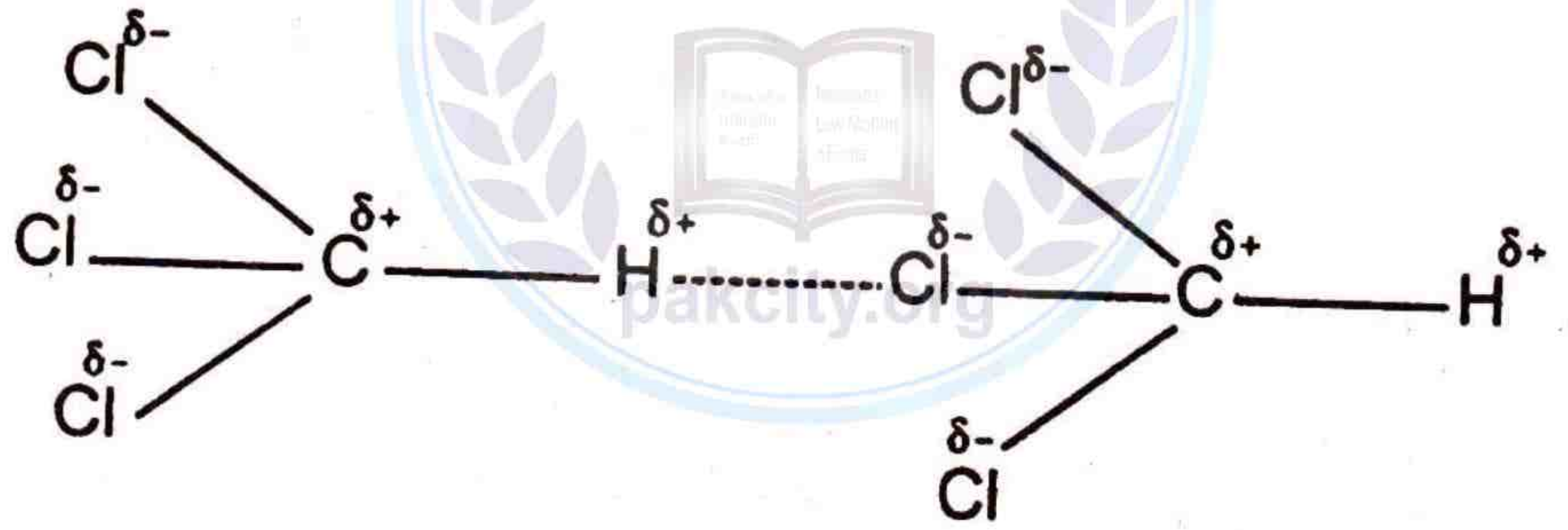
Dipole-dipole Forces

The positive end of one molecule attracts the negative end of the other molecule and these electrostatic forces of attraction are called dipole-dipole forces.

Example

In HCl molecule chlorine develops the partial negative charge and hydrogen develops the partial positive charge.





Dipole-dipole forces are approximately one percent as effective as a covalent bond.

Factors

The strength of these forces depends upon:

1. The electronegativity difference between the bonded atoms.
2. The distance between the molecules.

Strength of Forces in Different Phases

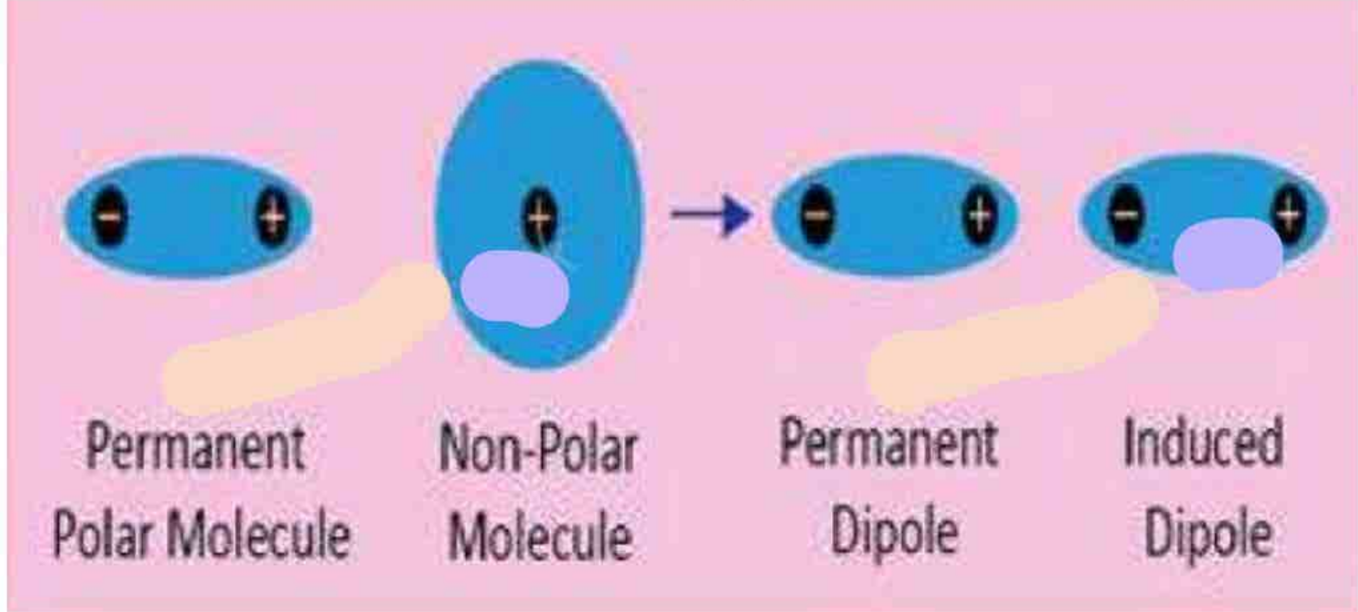
The distances between molecules in the gaseous phase are greater so these forces are very weak in this phase. In liquids these forces are reasonably strong.

Effect on Thermodynamic Parameters

Greater the strength of these dipole-dipole forces, greater are the values of thermodynamic parameters like melting points, boiling points, heats of vapourization and heats of sublimation.

Dipole-induced Dipole Forces

In a mixture of substances containing polar and non-polar molecules the positive end of the polar molecule attracts the mobile electrons of the nearby non-polar molecule. Polarity is induced in non-polar molecule, and both molecules become dipoles. These forces are called dipole-induced dipole forces or as Debye forces.



Instantaneous dipole-induced dipole forces or London Dispersion Forces

Definition

The momentary force of attraction created between instantaneous dipole and the induced dipole is called instantaneous dipole-induced dipole interaction or London force.

Explanation

A German physicist Fritz London in 1930 offered a simple explanation for these weak attractive forces between non-polar molecules.

Attractive forces among He atoms

1) **Generation of instantaneous dipole**

- In helium gas, the electrons of one atom influence the moving electrons of the other atom.
- Electrons repel each other and they tend to stay as far apart as possible.
- When the electrons of one atom come close to the electron of other atom, they are pushed away from each other.
- In this way, a temporary dipole is created in the atom.
- At any moment the electron density of the atom is no more symmetrical.
- It has more negative charge on one side than on the other.
- At that particular instant, the helium atom becomes a dipole.
- This is called **instantaneous dipole**.

2) **Generation of induced dipole**

- The instantaneous dipole then disturbs the electronic cloud of the other nearby atom.
- A dipole is induced in the second atom.
- This is called **induced dipole**.

Characteristics

- It is a very short-lived attraction because the electrons keep moving.
- London forces are present in all types of molecules whether polar or non-polar, but they are very' significant for non-polar molecules like Cl_2, H_2 and noble gases.

Factors Affecting the London Forces

1. Size of atom

When the size of the atom or molecule is large then the dispersion becomes easy and these forces become more prominent.

2. Atomic number

The atomic number increases down the group and the outermost electrons move away from the nuclei. The dispersion of the electronic clouds becomes more and more easy. So the polarizability of these atoms goes on increasing.

3. Polarizability

Polarizability is the quantitative measurement of the extent to which the electronic cloud can be polarized or distorted. This increased distortion of electronic cloud creates stronger London forces and hence the boiling points are increased down the group.

4. Boiling points of halogens

The boiling points of halogens in group VII-A also increase from fluorine to iodine. Fluorine is a gas and boils at -188.1 °C, while iodine is a solid at room temperature which boils at +184.4 °C. The polarizability of iodine molecule is much greater than that of fluorine.

5. Number of atoms

Greater the number of atoms in a molecule greater is its polarizability.

Example

The boiling points of C_2H_6 and C_6H_{14} are - 88.6 °C and 68.7 °C, respectively. The reason is that longer molecules have more places along its length where they can be attracted to other molecules. With the increasing molecular mass of these hydrocarbons, they change from gaseous to liquid and then to solids.

Hydrogen Bonding

Definition

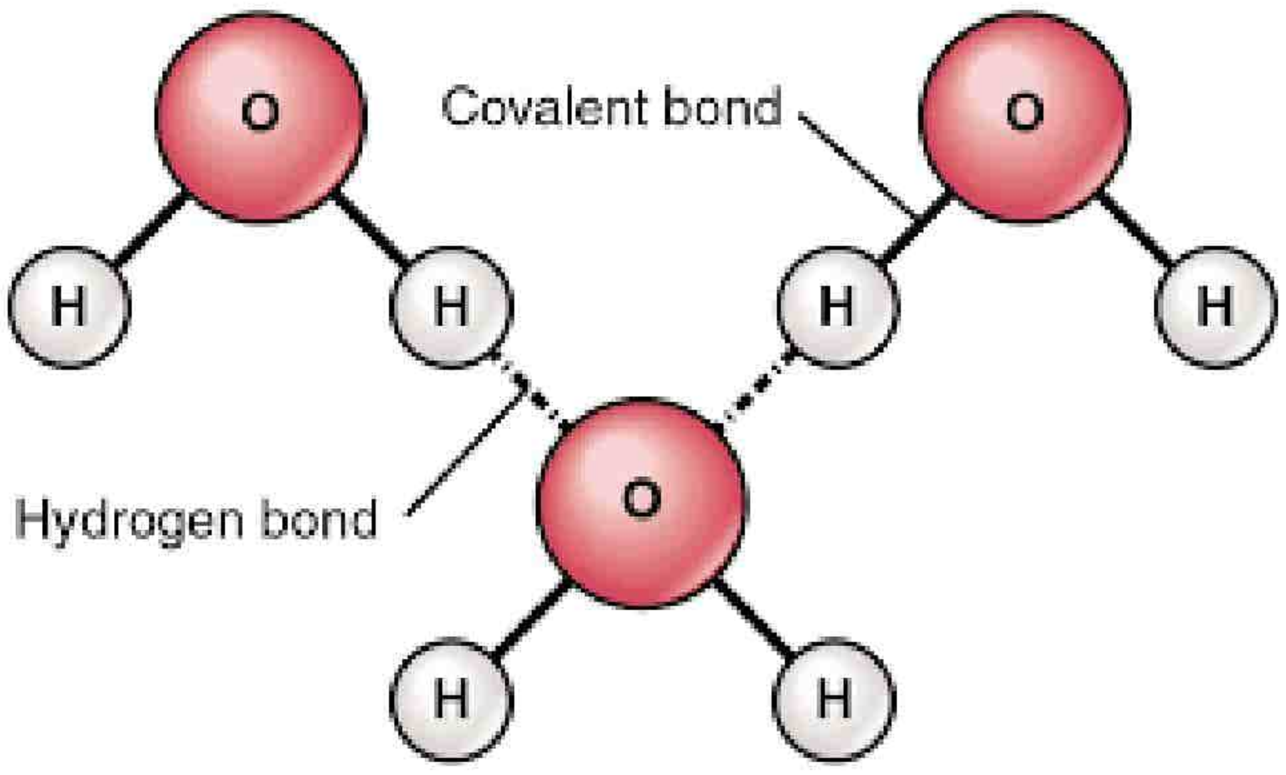
Hydrogen bonding is the electrostatic force of attraction between a highly electronegative atom and partial positively charged hydrogen atom.

Example
Hydrogen bonding in the molecule of water.

Explanation
- Oxygen is more electronegative element as compared to hydrogen, so water is a polar molecule.
- There is dipole-dipole interactions between partial positively charged hydrogen atoms and partial negatively charged oxygen atoms.
- Oxygen atom has two lone pairs.
- Hydrogen has sufficient partial positive charge.
- Both the hydrogen atoms of water molecule create strong electrical field due to their small sizes.
- The oxygen atom of the other molecule links to form a coordinate covalent bond with hydrogen using one of its lone pairs of electrons.

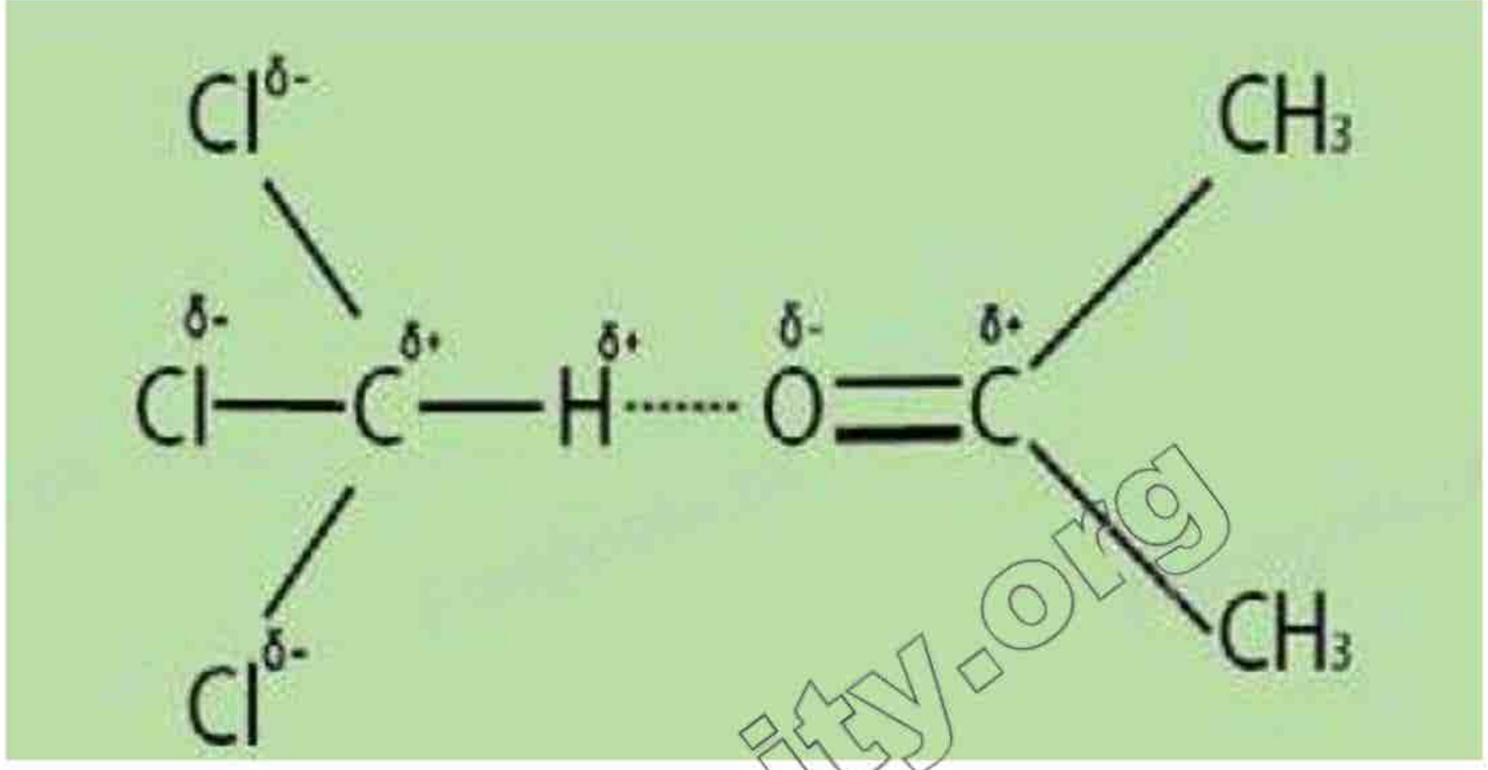
Electronegative atoms creating hydrogen bonding
- Fluorine
- Oxygen
- Nitrogen
- Rarely chlorine

Strength of hydrogen bond
The strength of hydrogen bond is generally twenty times less than that of a covalent bond.

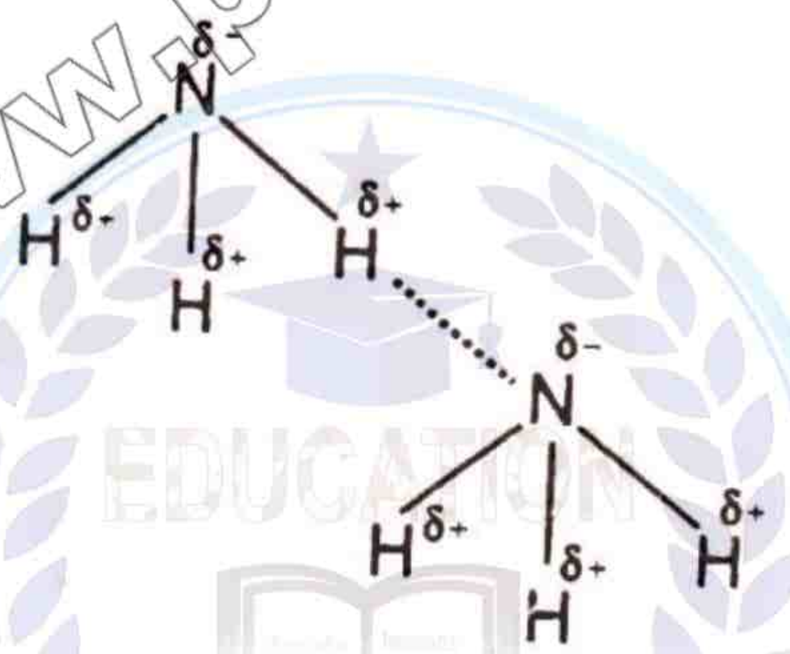


Other Examples of Hydrogen Bonding

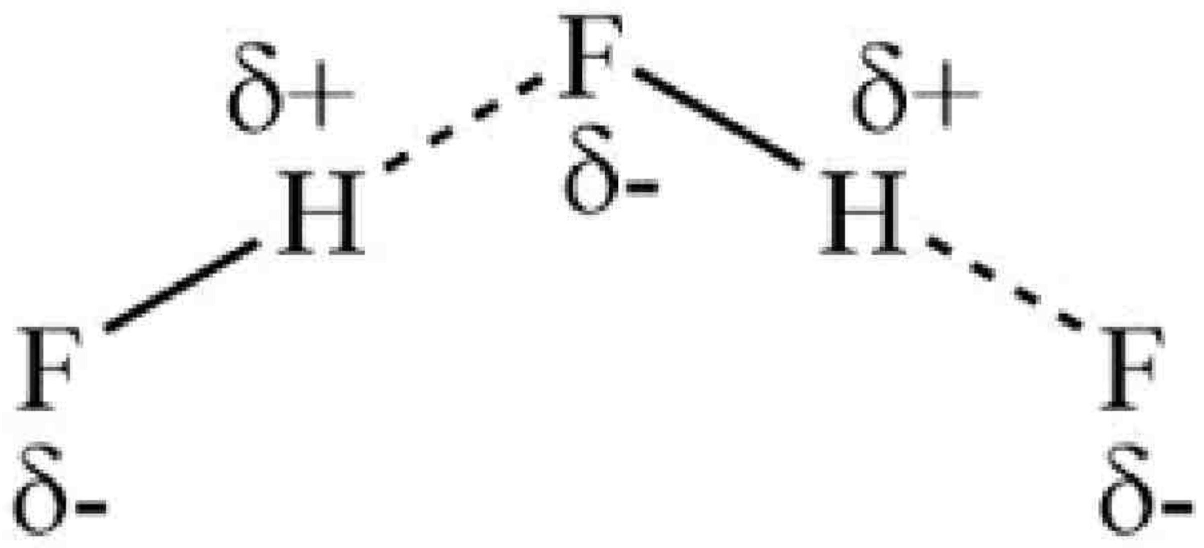
Hydrogen bonding between acetone and chloroform



Hydrogen bonding in ammonia and hydrofluoric acid



The molecules of HF join with each other in a zig-zag manner.



The bond angle is 120 ^{o}

Low Acidic Strength of HF

The low acidic strength of HF molecule as compared to HCl, HBr and HI is due to the strong hydrogen bonding, because the partial positively charged hydrogen is entrapped between two highly electronegative atoms.

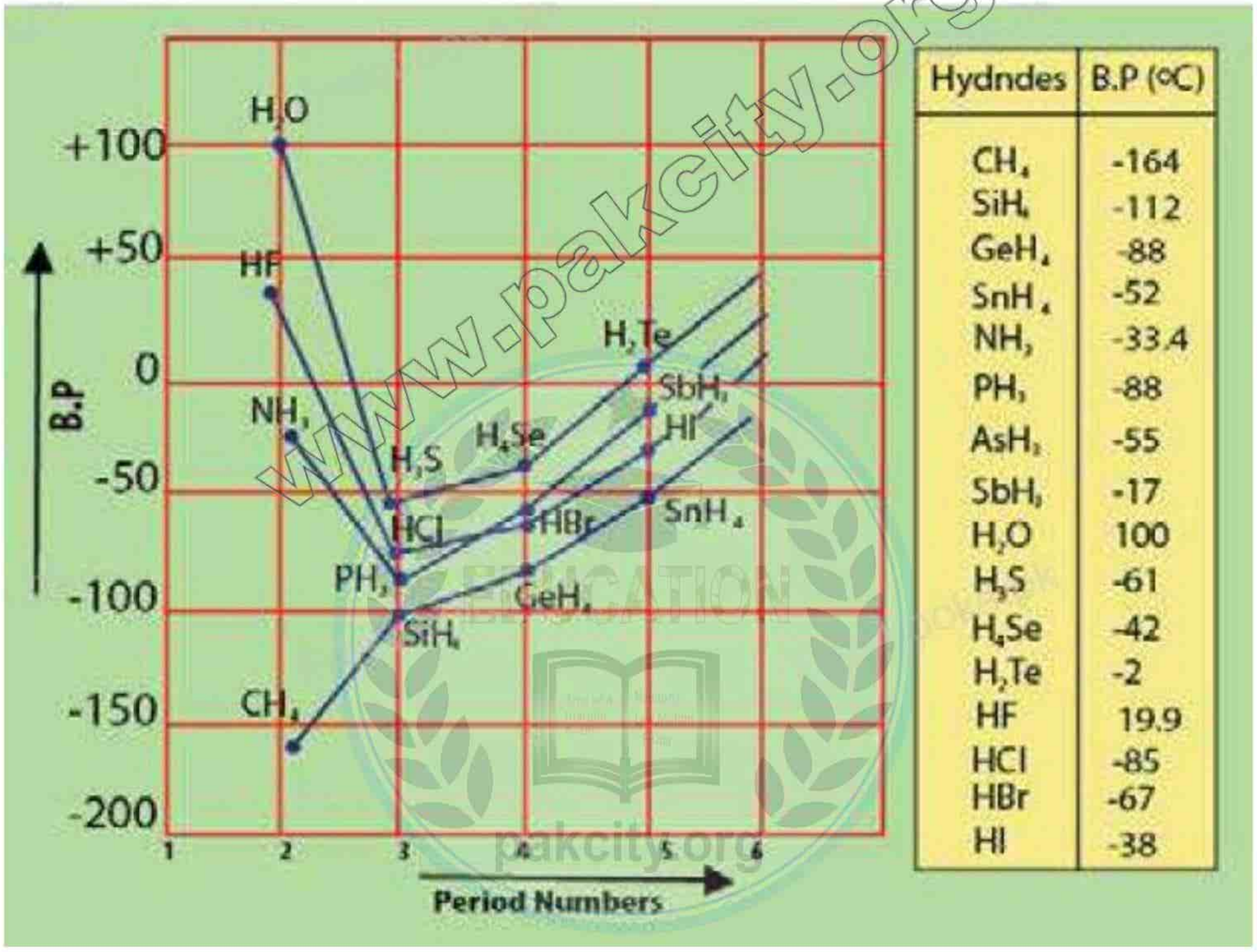
Properties and Application of Compounds containing Hydrogen Bonding

1. Thermodynamic Properties of Covalent Hydrides

Hydrogen bonding influences the physical properties like melting and boiling points.

Comparison of the physical properties of hydrides of IV-A, V-A, VI-A and VII-A

For comparison the graphs are plotted between the period number of the periodic table on x-axis and boiling points in kelvin on y-axis.



Hydndes	B.P (°C)
CH_4	-164
SiH_4	-112
GeH_4	-88
SnH_4	-52
NH_3	-33.4
PH_3	-88
AsH_3	-55
SbH_3	-17
H_2O	100
H_2S	-61
H_2Se	-42
H_2Te	-2
HF	19.9
HCl	-85
HBr	-67
HI	-38

Conclusions from the graph

Boiling point of hydrides of group IV-A

Hydrides of group IV-A have low boiling points as compared to those of group V-A, VI-A, VII-A. **Reason:** These elements are least electronegative. Methane has the lowest boiling point because it is a very small molecule and its polarizability is the least.

Boiling point of hydrides of V-A, VI-A, VII-A

The hydrides of group V-A, VI-A, VII-A NH_3, H_2O and HF show maximum boiling points in the respective series. **Reason:** The reason is the enhanced electronegative character of N, O and F.

- Water is liquid at room temperature, but H_2S and H_2Se are gases.

- Fluorine is more electronegative than oxygen. So, H-bonding in HF is expected to be stronger than that in water and the boiling point of HF higher than that of H_2O. However, it is lower and the reason is that the fluorine atom can make only one hydrogen bond per molecule.
- Water can form two hydrogen bonds per molecule, as it has two hydrogen atoms and two lone pairs on oxygen atom.
- Ammonia can form only one hydrogen bond per molecule as it has only one lone pair.

Boiling point of halogen acids

The boiling point of HBr is slightly higher than that of HCl. It means that chlorine is electronegative enough to form a hydrogen bond. It seems that HCl has a strong dipole-dipole interaction but it is a border line case.

Boiling point of hydrides of fourth period

The hydrides of fourth period GeH_4, AsH_3, H_2Se, HBr show greater boiling points than those of third period due to greater size and enhanced polarizabilities.

2. Solubility of Hydrogen-Bonded Molecules

Example 1

Water is the best example of H-bonded system. Similarly ethyl alcohol (C_2H_5OH) also has the tendency to form hydrogen bonds. Ethyl alcohol can dissolve in water because both can form hydrogen bonds with each other.

Example 2

Small-sized carboxylic acids are soluble in water. Hydrocarbons are not soluble in water because they are non-polar compounds and there is no hydrogen bonding between water and hydrocarbon molecules.

3. Structure of Ice

1. The molecules of water have tetrahedral structure.
2. Two lone pairs of electrons on oxygen atom occupy two corners of the tetrahedron.
3. In the liquid state, water molecules are extensively associated with each other and these associations break and are reformed because the molecules of water are mobile.
4. When the temperature of water is decreased and ice is formed then the molecules become more regular and this regularity extends throughout the whole structure.
5. Empty spaces are created in the structure.
6. When water freezes it occupies 9% more space and its density decreases.
7. The result is that ice floats on water. The structure of ice is just like that of a diamond being tetrahedral.

Hydrogen Bonding and Life Sustenance under frozen water

The lower density of ice than liquid water at 0 ºC causes water in ponds and lakes to freeze from surface to the downward direction. The less dense water below 4 ºC stays on the top of slightly warm water underneath and ultimately it freezes. This layer of ice insulates the water underneath and fish and plants survive under this blanket of ice.

4. Cleansing Action of Soaps and Detergents

Soaps and detergents perform the cleansing action because:

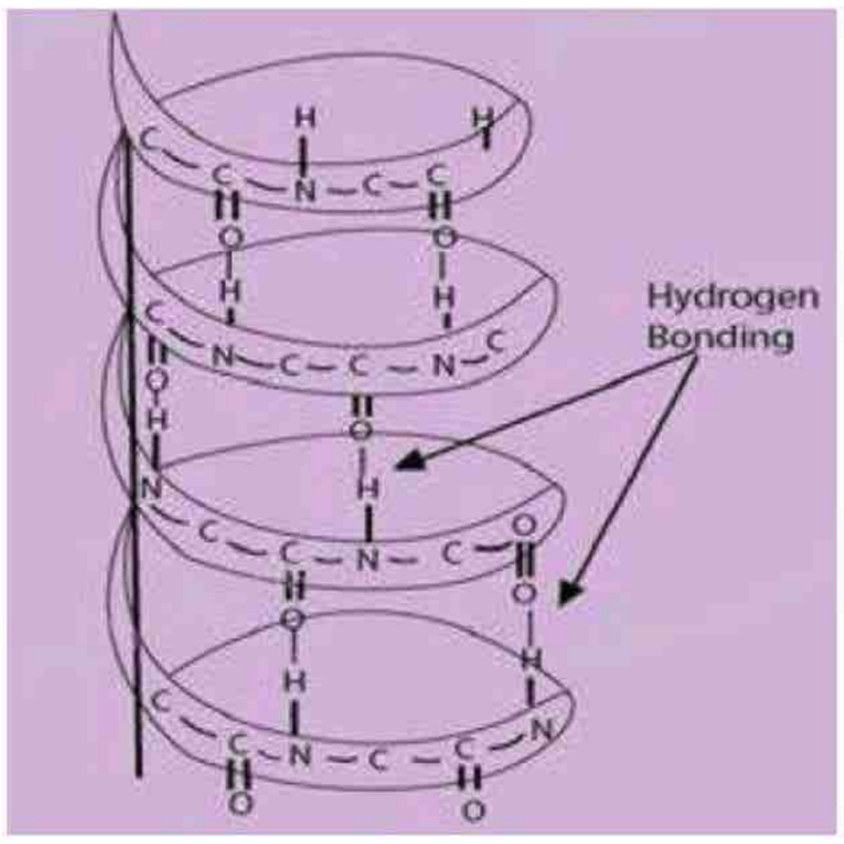
1. The polar part of their molecules are water soluble due to hydrogen-bonding
2. The non-polar parts remain outside water, because they are alkyl or benzyl portions and are insoluble in water.

5. Hydrogen Bonding in Biological Compounds and Food Materials

Hydrogen bonding exists in the molecules of living system.

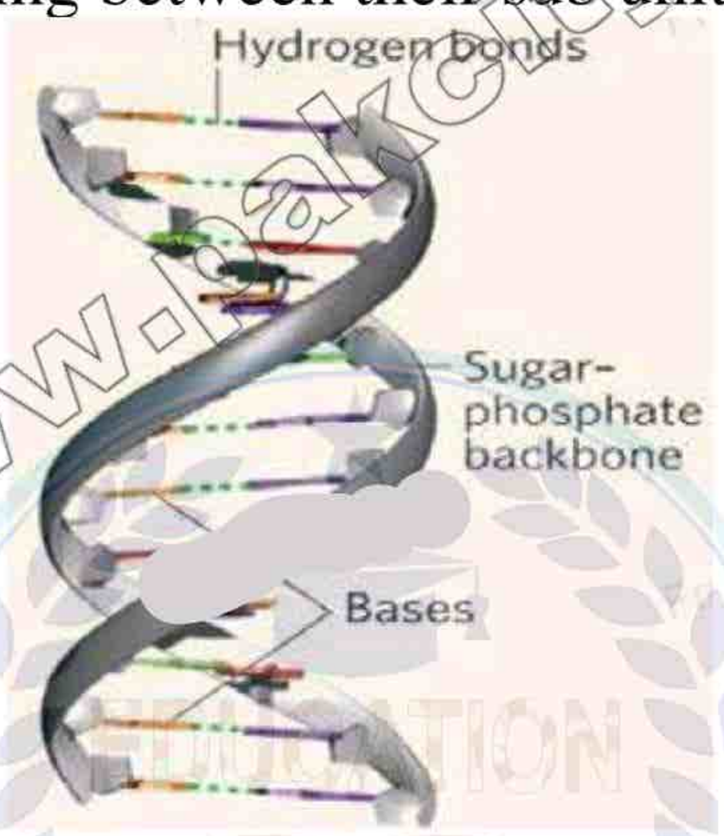
Structure of Proteins

1. Hair, silk and muscle fibres consist of long chains of amino acids.
2. These chains are coiled about one another into a spiral called a helix.
3. Helix may be right handed or left handed.
4. In right handed helix the groups like >N H and > C = O are vertically adjacent to one another linked by hydrogen bonds.
5. These H-bonds link one spiral to the other. X-ray analysis confirms 27 amino acid units on average for each turn of the helix.



Structure of DNA

1. Deoxyribonucleic acid (DNA) has two spiral chains.
2. These are coiled about each other on a common axis.
3. They give a double helix 18-20 Å in diameter.
4. They are linked together by H-bonding between their sub units.



Hydrogen Bonding in Carbohydrates

The food materials like carbohydrates include glucose, fructose and sucrose. They all have -OH groups in them which are responsible for hydrogen bonding in them.

6. Hydrogen Bonding in Paints, Dyes and Textile Materials

H-bonding in paints and dyes

The adhesive action of paints and dyes is due to hydrogen bonding.

H-bonding in glue and honey

Hydrogen bonding makes glue and honey as sticky substances.

H-bonding in cotton, silk or synthetic fibres

Hydrogen bonding is of vital importance in the thread making materials. This hydrogen bonding is responsible for their rigidity and tensile strength.

Manometric Method

Vapour Pressure

Definition

The vapour pressure of a liquid is a pressure exerted by the vapours of the liquid in equilibrium with the liquid at a given temperature.

$$\text{Liquid} \rightleftharpoons \text{Vapour}$$

Manometric method is an accurate method for determining vapour pressure.

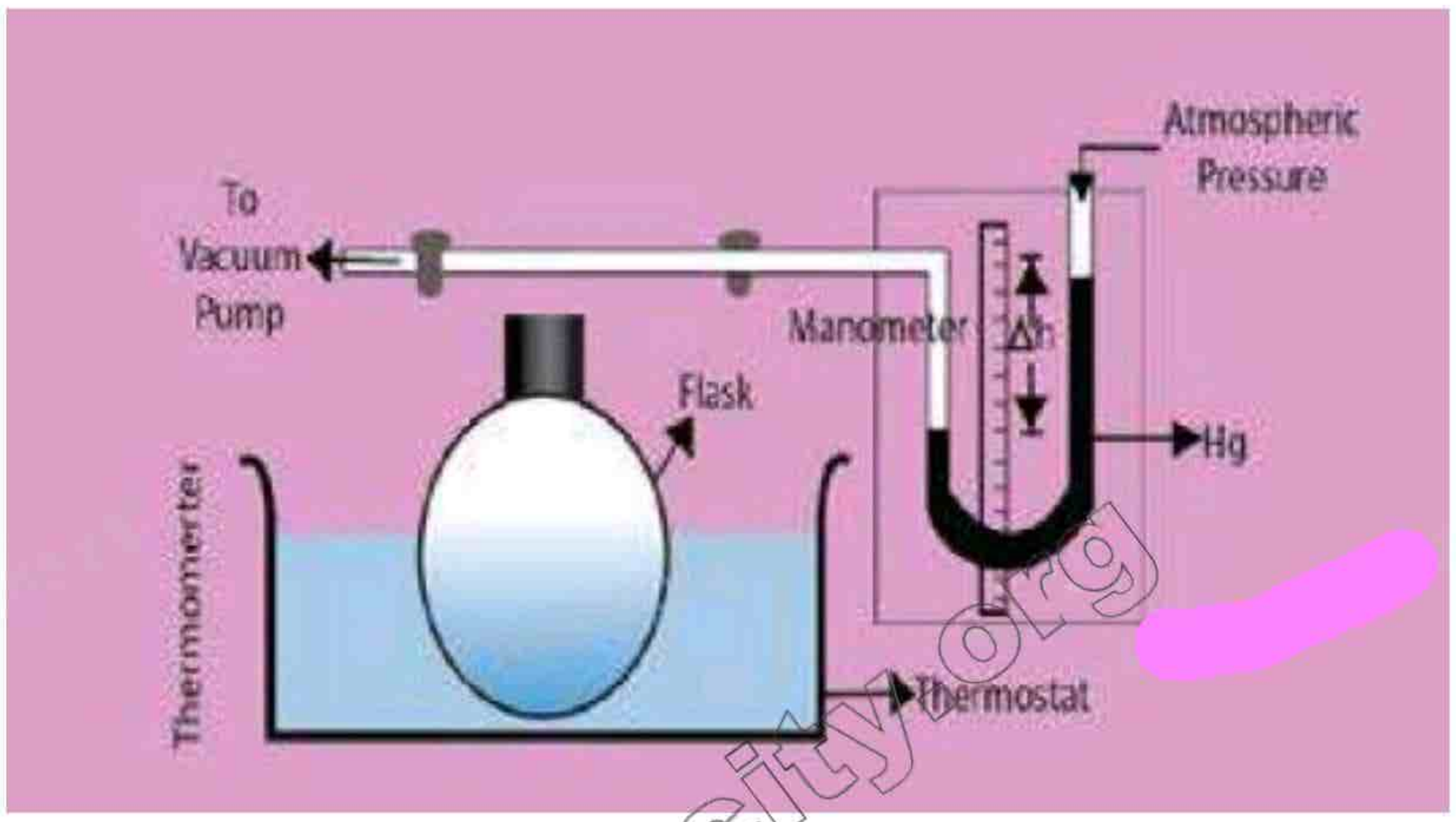
Assembly

Liquid in flask

The liquid whose vapour pressure is to be determined is taken in a flask placed in a thermostat.

Manometer and Vacuum pump

One end of the tube from the flask is connected to a manometer and the other end is connected to a vacuum pump.



Working

Freezing of liquid

1. The liquid is frozen with the help of a freezing mixture and the space above the liquid is evacuated.
2. The air is removed from the surface of the liquid along with the vapours of that liquid.
3. The frozen liquid is then melted to release any entrapped air.
4. Liquid is again frozen and released air is evacuated.
5. This process is repeated many times till almost all the air is removed.

Warming of liquid

6. The liquid is warmed in the thermostat to that temperature at which its vapour pressure in the flask is to be determined.

Determination of vapour pressure

7. Difference in the heights of the columns of Hg in liquid determines the vapour pressure of the liquid.
8. The column of mercury in the manometer facing the vapours of the liquid is depressed.
9. The other column which faces the atmospheric pressure rises.
10. The pressure on the surface of the liquid in the flask is equal to the sum of the atmospheric pressure and the vapour pressure of liquid.
11. The column of manometer facing the liquid is more depressed than facing the atmosphere.

Equation used

$$P = Pa + \Delta h$$

P= vapour pressure of the liquid at one atm pressure.

Pa=Atmospheric pressure

Δh= Difference in the heights of the mercury levels in the two limbs of the manometer giving us the vapour pressure of liquid

Liquid Crystals

Historical Background

In 1888, Frederick Reinitzer, an Austrian botanist discovered a universal property. He was studying an organic compound cholesteryl benzoate. This compound turns milky liquid at 145°C and becomes a clear liquid at 179°C. When the substance is cooled, the reverse process occurs. This turbid liquid phase was called liquid crystal.

Definition

Liquid crystalline state exists between two temperatures i.e. melting temperature and clearing temperature.

$$\text{Crystal} \rightleftharpoons \text{Liquid crystal} \rightleftharpoons \text{Liquid}$$

Properties

1. Behave as liquids

Liquid crystals can flow as liquids. They have the properties like liquids as surface tension, viscosity, etc.

2. Behave as solids

Liquid crystals have some degree of order. It means that they resemble crystals in certain properties like optical ones.

3. Structure

Those substances which make the liquid crystals are often composed of long rod like molecules. In the normal liquid phase, these molecules are oriented in random directions. In liquid crystalline phase, they develop some ordering of molecules. Depending upon the nature of ordering, liquid crystals can be divided into nematic, smectic and cholesteric.

4. Anisotropic

The properties of liquid crystals are intermediate between those of crystals and isotropic liquids. They have the fluidity of the liquids and the optical properties of the crystals. A crystalline solid may be isotropic or anisotropic, but liquid crystals are always anistropic.

Uses of Liquid Crystals

1. As temperature sensors

Liquid crystals can diffract light. When one of the wavelengths of white light is reflected from a liquid crystal it appears coloured. With temperature change the distances between the layers of the molecules of liquid crystals change. The colour of the reflected light changes accordingly. Liquid crystals can he used as temperature sensors.

2. Detection of potential failure in electrical circuits

Liquid crystals are used to find the point of potential failure in electrical circuits.

3. In room thermometers

Room thermometers also contain liquid crystals with a suitable temperature range. As the temperature changes, figures show up in different colours.

4. Medical applications

Blockages in veins

Liquid crystalline substances are used to locate the veins, arteries, infections and tumors. The reason is that these parts of the body are warmer than the surrounding tissues. Specialists can use the techniques of skin thermography to detect blockages in veins and arteries.

Early diagnosis of breast cancer

When a layer of liquid crystal is painted on the surface of the breast, a tumor shows up as a hot area which is coloured blue. This technique has been successful in the early diagnosis of breast cancer.

4. Display of electrical devices

Liquid crystals are used in the display of electrical devices such as digital watches, calculators and laptop computers. These devices operate due to the fact that temperature, pressure and electro-magnetic fields easily affect the weak bonds, which hold molecules together in liquid crystals.

5. Use in chromatography

In chromatographic separations, liquid crystals are used as solvents.

6. Oscillography and TV display

Oscillographic and TV displays also use liquid crystal screens.

Classification of Solids

Ionic Solids

Definition

Crystalline solids in which the particles forming the crystal are positively and negatively charged ions are called ionic solids. These ions are held together by strong electrostatic forces of attraction. These attractive forces are also called ionic bonds.

Examples

NaCl, KBr etc.

Properties of Ionic Solids

1. Well-defined geometrical pattern

The cations and anions are arranged in a well defined geometrical pattern, so they are crystalline solids at room temperature.

2. Stable compounds

Under ordinary conditions of temperature and pressure they never exist in the form of liquids or gases. Ionic crystals are very stable compounds.

3. High melting and boiling points

Very high energy is required to separate the cations and anions from each other against the forces of attraction. That is why ionic crystals are very hard, have low volatility and high melting and boiling points.

4. Non-existence as neutral molecules

Ionic solids do not exist as individual neutral independent molecules.

5. Non-directional bonds

The cations and anions attract each other and these forces are non-directional.

6. Close packing

The close packing of the ions enables them to occupy minimum space.

7. Systematic arrangement of ions

A crystal lattice is developed when the ions arrange themselves systematically in an alternate manner. The structure of the ionic crystals depends upon the radius ratio of cations and anions. For example, NaCl and CsF have the same geometry because the radius ratio in both the cases is the same.

8. Formula mass

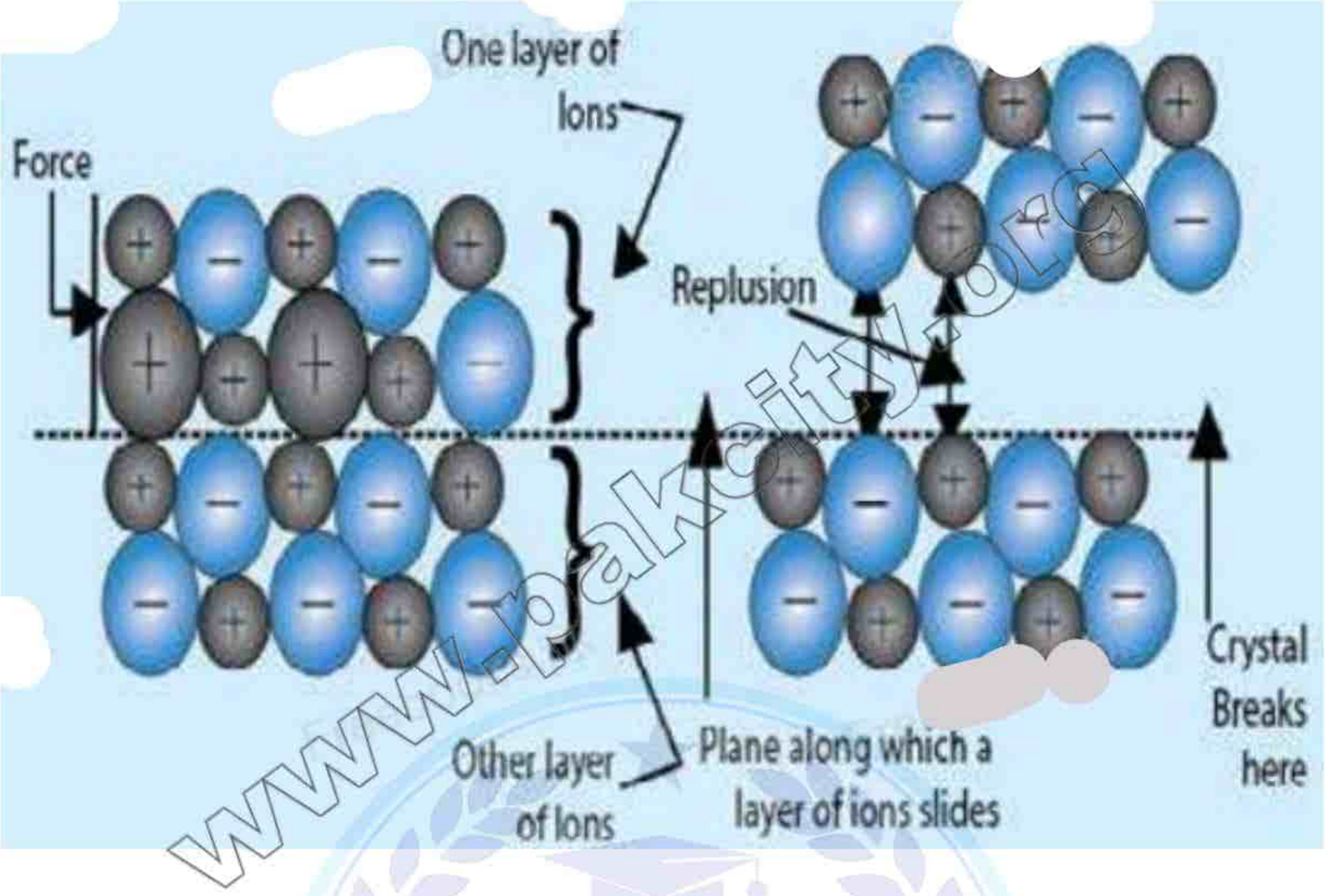
The term formula mass for the ionic substances is used and not the molecular mass, because they do not exist in the form of molecules.

9. Non-conductors in solid state

Ionic crystals do not conduct electricity in the solid state, because on account of electrostatic force existing between them the cations and anions remain tightly held together and hence occupy fixed positions. Ionic crystals conduct electricity when they are in solution or in the molten state. In both cases ions become free.

10. Brittleness

Ionic crystals are highly brittle because ionic solids are composed of parallel layers which contain cations and anions in alternate positions, so that the opposite ions in the various parallel layers lie over each other. When an external force is applied, one layer of the ions slides a bit over the other layer along a plane. In this way the like ions come in front of each other and repel. This causes brittleness.



11. High density

Ionic solids are mostly of high density due to close packing of ions.

12. Ionic reactions

Such compounds having the ionic crystals give ionic reactions in polar solvents and these are very fast reactions.

13. Isomorphism and polymorphism

The properties like isomorphism and polymorphism are also associated with the ionic crystals.

Structure of Sodium Chloride

Structure and size of ions

1. The structure of ionic crystals depends upon the structure and the size of their ions.
2. Each ion is surrounded by a certain number of ions of opposite charge.

3. In the structure of NaCl each Na^+ ion is surrounded by six chloride ions.
4. It is clear that Na^+ has ten electrons while Cl^- has total 18 electrons.
5. The size of the Cl^- is bigger than that of Na^+.

Distance between ions

1. The distance between two nearest ions of the same kind i.e., Cl^- ions is 5.63 ^oA .
2. The distance between two adjacent ions of different kind is $5.63/2 = 2.815$ ^oA.

Coordination number

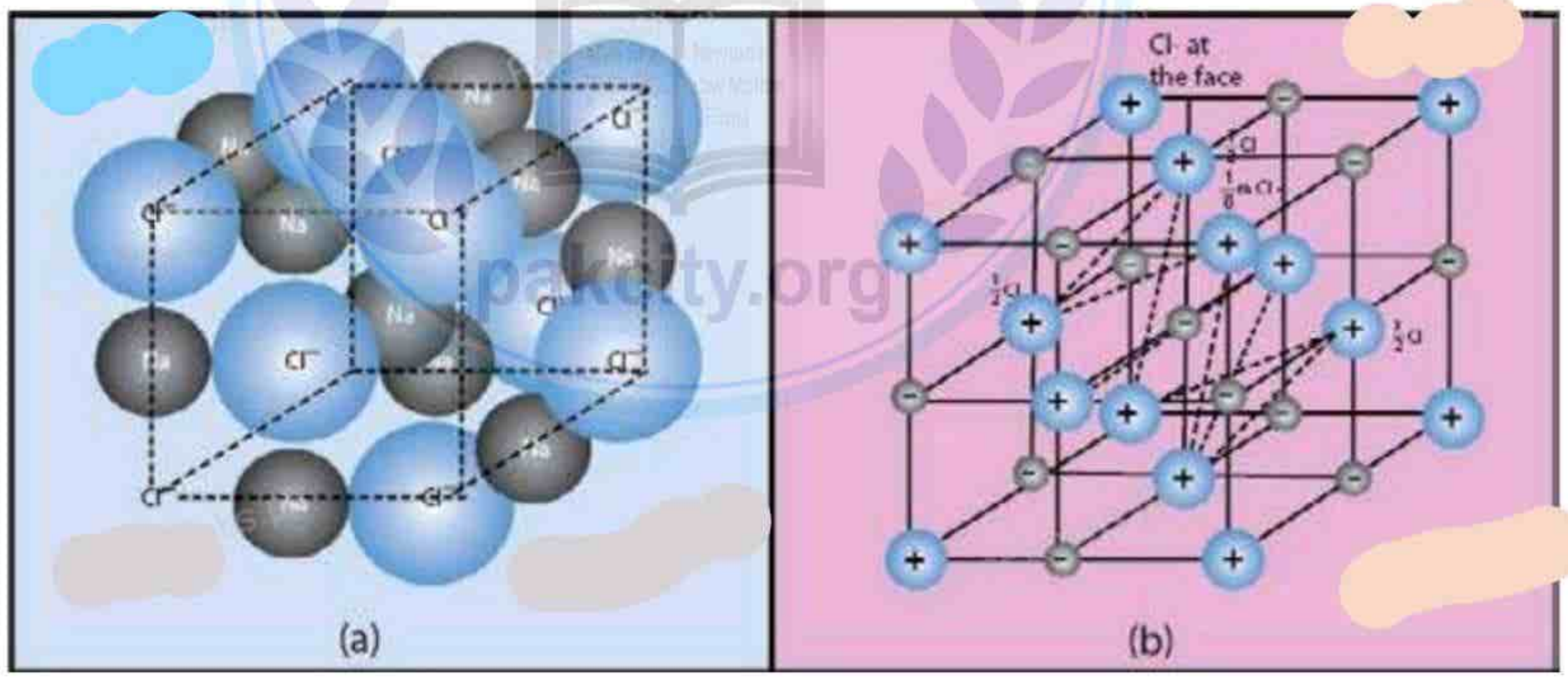
1. Each Na^+ is surrounded by six Cl^- placed at the corners of a regular octahedron.
2. The coordination number of each Na^+ is six.
3. Each Cl^- is also surrounded by six Na^+.

Formula unit

1. Na^+ and $C1^-$ are not connected to one another by pairs because all six Cl^- ions are at the same distance away from one Na^+.
2. Independent molecules of NaCl do exist in the vapour phase.
3. In solid NaCl there are no independent molecules of NaCl.
4. Sodium chloride is said to have formula unit of NaCl.

Chlorides at the corners and face centres of cube

1. There are eight Cl^- at the comers of the cube, and each is being shared amongst eight cubes.
2. l/8th part of each Cl^- ion is considered for this unit cell.
3. One complete Cl^- is contributed by eight corners.
4. Six chloride ions are present at the face centres and each is being shared between two cells.
5. Per unit cell there are $8/8 + 6/2 = 4$ Cl^- ions.
6. Justification of 4 Na^+ can be given, if a unit cell has 8 Na^+ at eight corners and 6 Na^+ at faces.
7. There are equal number of Na^+ ions
8. Four NaCl units are present per unit cell.



Covalent Solids

Definition

Covalent solids are also called atomic solids, because they are composed of neutral atoms of the same or of different elements. These atoms are held together by covalent bonds.

Types of covalent solids

Covalent solids are of two types.

Giant molecules

(i) When the covalent bonds join to form giant molecules like diamond, silicon carbide or aluminium nitride.

Layered structures

(ii) When atoms join to form the covalent bonds and separate layers are produced like that of graphite, cadmium iodide and boron nitride.

Properties of covalent solids

1. Three dimensional structure

The bonding in covalent crystals extend in three dimensions. They contain a network of atoms.

2. Directional bond and open structure

The valencies of atoms are directed in definite directions, so the packing of atoms in these crystals is looser than those of ionic and metallic crystals. Covalent crystals have open structure.

3. High melting and boiling points

These crystals are very hard and considerable amount of energy is required to break them. They have high melting points and their volatility is very low.

4. Bad conductors of electricity

Due to the absence of free electrons and ions they are bad conductors of electricity. Graphite has a layered structure and the electrons are available in between the layers. These electrons are delocalized and conductivity becomes possible. Graphite is not a conductor perpendicular to the layers.

5. Solubility in non-polar solvents

These solids are insoluble in polar solvents like water but they are readily soluble in non-polar solvents like benzene and carbon tetrachloride.

6. Rate of chemical reaction of giant molecules

The covalent crystals having giant molecules like diamond and silicon carbide are insoluble in all the solvents. Because of their big size, they do not interact with the solvent molecules. The chemical reactions of such crystalline solids are very slow.

Structure of diamond

sp^3 hybridization

The four atomic orbitals (one 2s and three 2p) undergo sp^3 hybridization to give four sp^3 hybridized orbitals. They are directed in space along the four corners of a tetrahedron.

sp^3-sp^3 overlap

This is the unit cell of diamond and a large number of such unit cells undergo sp^3-sp^3 overlapping to form a huge structure. Each carbon atom is linked with four other carbon atoms. The bonds between carbon atoms are covalent which run through the crystal in three-dimensions.

Bond angles and bond lengths

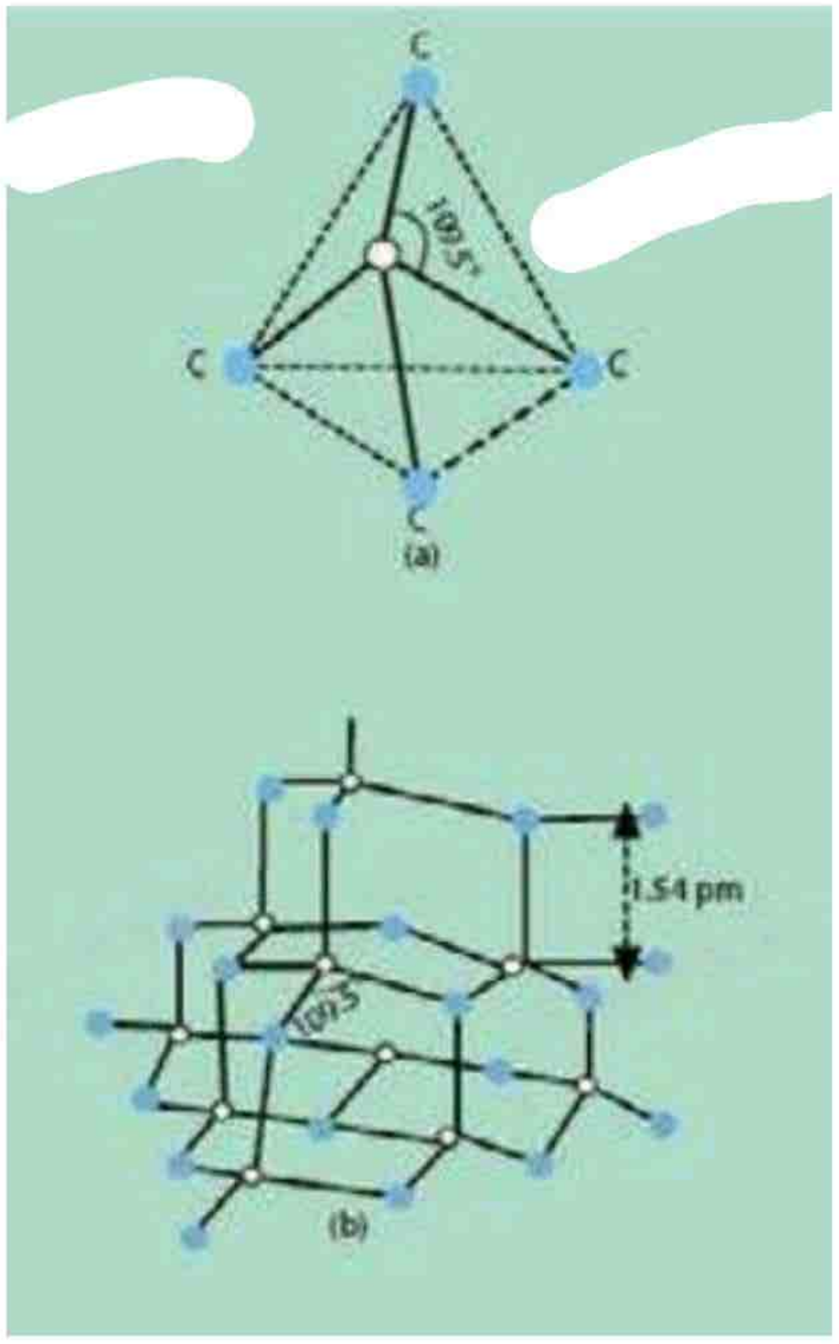
All the bond angles are 109.5° and the bond lengths are 154 pm.

Continuous three dimensional structure

The whole lattice is continuous and because of the continuity of C-C covalent bonding the entire diamond crystal behaves as a huge or giant three-dimensional carbon molecule.

Face centred-cubic structure

The overall structure of diamond looks face centred-cubic.



Molecular Solids

Definition

Those solid substances in which the particles forming the crystals are polar or non-polar molecules or atoms are called molecular solids.

Types of intermolecular forces

Two types of intermolecular forces hold them together.

(i) Dipole-dipole interactions.

(ii) van der Waals forces.

These intermolecular forces are much weaker than those in ionic crystals and covalent crystals.

Polar molecules

Ice and sugar are the best examples of crystals having polar molecules.

Non-polar molecules

Iodine, sulphur, phosphorus, carbon dioxide and solidified noble gases form the non-polar molecules.

Melting and boiling points

Polar molecular solids have usually higher melting and boiling points as compared to non-polar molecular solids.

Properties of the molecular solids

1. Regular arrangement of atoms

X-ray analysis has shown the regular arrangements of atoms in constituent molecules of these solids and we get the exact positions of all the atoms.

2. Weak forces

The forces, which hold the molecules together in molecular crystals, are very weak so they are soft and easily compressible.

3. Volatile and low melting and boiling points

They are mostly volatile and have low melting and boiling points.

4. Bad conductors of electricity

They are bad conductors of electricity.

5. Low densities

They have low densities.

6. Transparency to light

They are sometimes transparent to light.

7. Solubility

Polar molecular crystals are mostly soluble in polar solvents, while non-polar molecular crystals are usually soluble in non-polar solvents.

Structure of iodine

Layer lattice

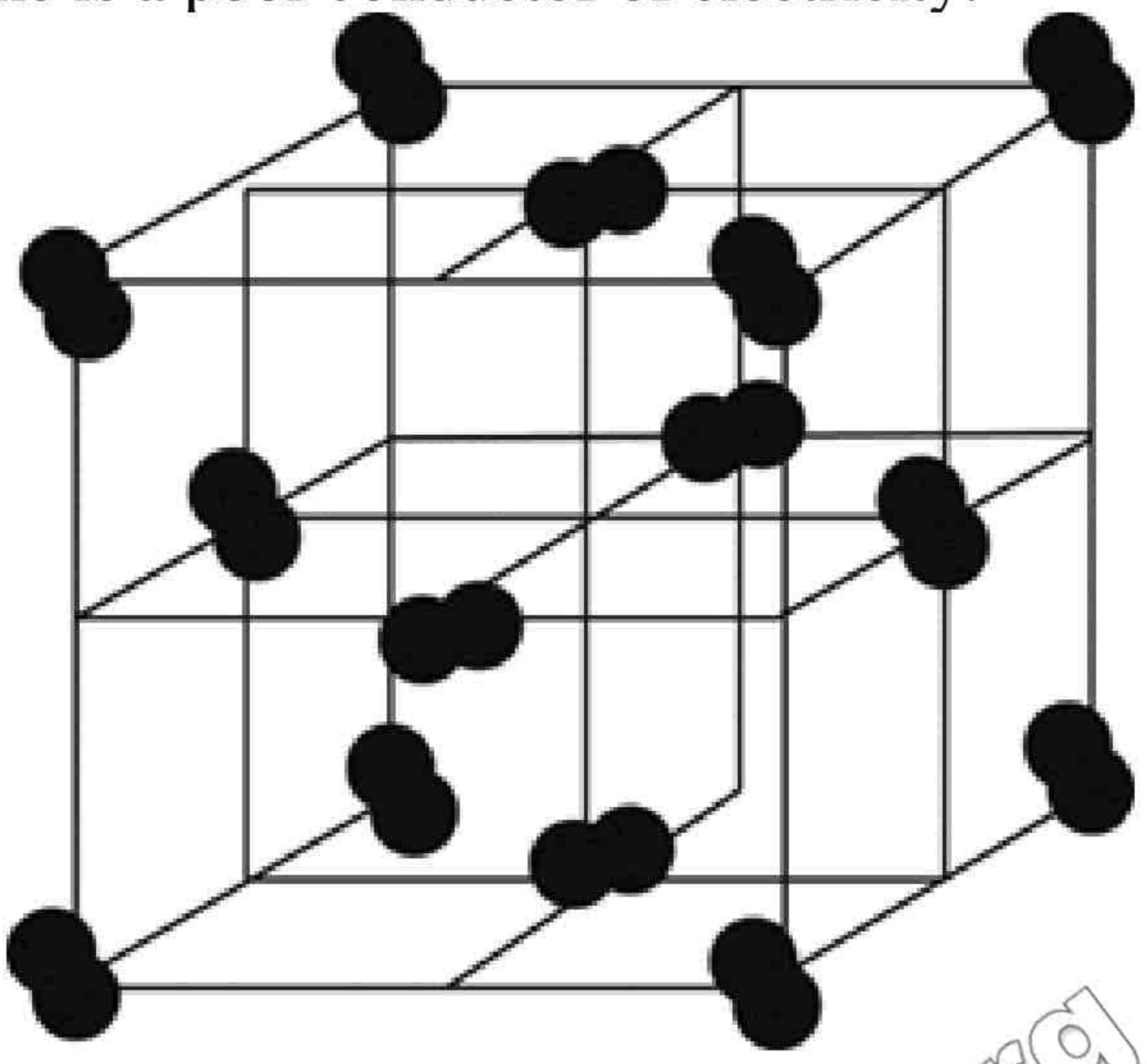
In the solid state the molecules of iodine align in the form of layer lattice.

I-I bond distance

I-I bond distance is 271.5 pm and it is appreciably longer than in gaseous iodine (266.6 pm).

Poor conductor of electricity

As expected from its structure, iodine is a poor conductor of electricity.



Metallic Solids

Definition

The force which binds a metal cation to a number of electrons within its sphere of influence is known as metallic bond.

Theories of Metallic Bonding

Electron pool or electron gas theory

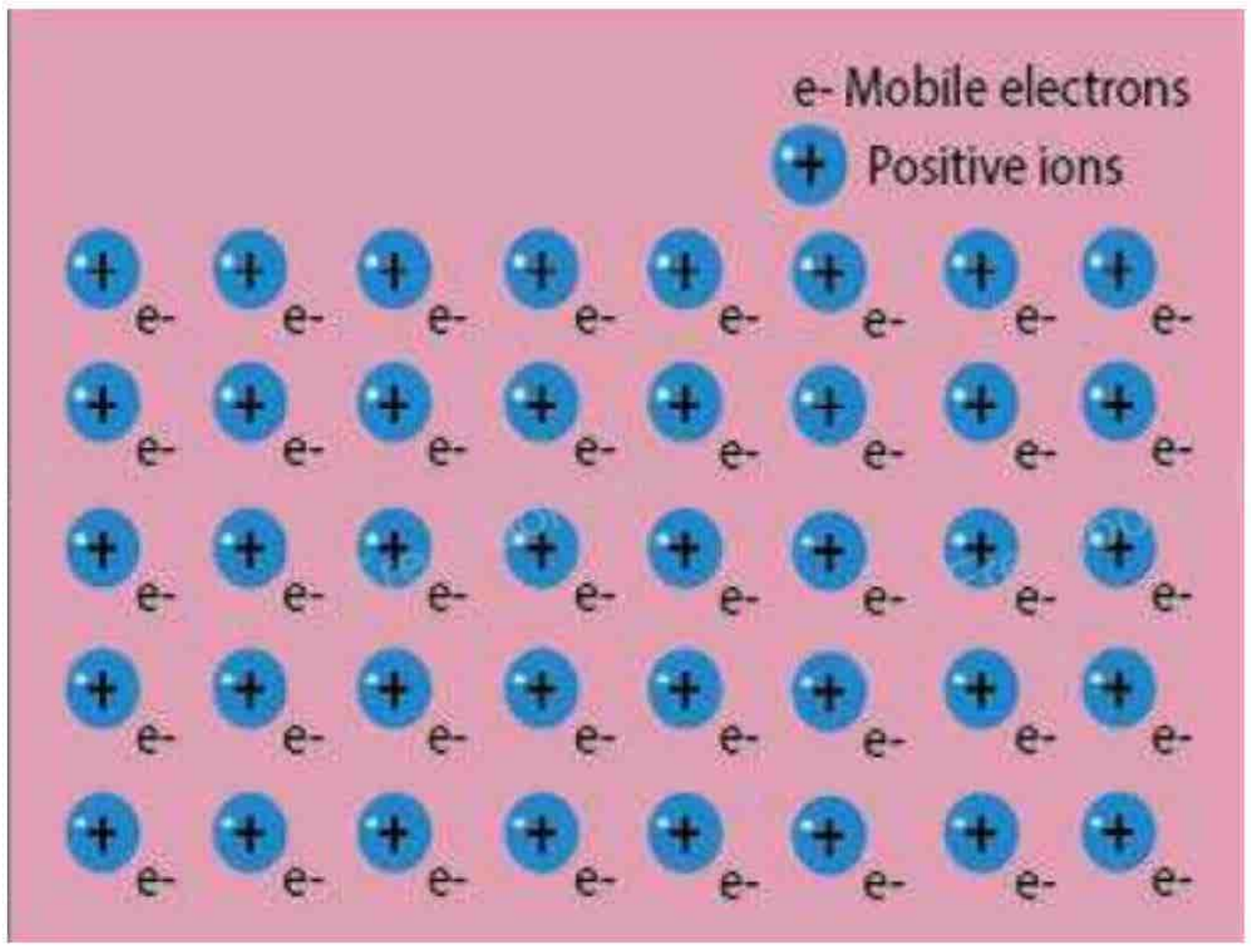
The first theory of metallic bonding is called electron pool or electron gas theory. This theory was **proposed by Drude and extended by Loren (1923)**.

According to this theory:

Each atom in a metal crystal
loses all of its valence electrons. These
valence electrons form a pool or a gas.
The positively charged metal ions are
believed to be held together by electron
pool or gas.

Explanation

1. The positively charged ions occupy definite positions at measurable distances from each other in the crystal lattice.
2. Valence electrons are not attached to any individual ion or pair of ions rather belong to the crystal as a whole.
3. These electrons are free to move about from one part of the crystal to the other.



Valence Bond Theory

L. Pauling explained the metallic bond according to valence bond theory.

Statement

According to this theory:

The metallic bond is treated essentially as covalent in character. It is assumed that the covalent bonds are not localized but are highly delocalized in metal structure.

Molecular Orbital Theory

Molecular orbital theory was applied to explain the characteristics of metallic solids.

Statement

According to this theory:

It is assumed that the electrons in the completely filled orbitals are essentially localized, while atomic orbitals containing the valence electrons interact or overlap to form a set of delocalized orbitals.

Explanation

1. The delocalized orbitals are the molecular orbitals which extend over the entire crystal lattice.
2. Such a combination of atomic orbitals produce as a large number of closely spaced states.
3. These states of energy are known as **bands of energy**.
4. It is called a **band theory**.
5. The energy gap between two bands determines the properties of the metallic solids.

Properties of metallic solids

1. Electrical conduction

Metals are good conductor of electricity. When electric field is applied between two ends of a metal then the mobile electrons begin to move towards the positive pole and the new electrons from the negative pole take their place.

2. Effect of temperature on electrical conduction

The electrical conductivity of metals decreases with the increase in temperature. With the increase in temperature the positive metal ions also begin to oscillate and the motion hinders the free movement of mobile electrons between the positive ions. This hindrance decreases the electrical conductivity.

3. Thermal conductivity

When a piece of metal is heated at one end, the mobile electrons at this end absorb heat energy and move towards the cooler end. During the process they collide with adjacent electrons and transfer their heat energy to them.

4. Metallic luster

Whenever the metals are freshly cut, most of them possess metallic luster. When light falls on the metallic surface, the incident light collides with the mobile electrons and they are excited. These electrons when de-excited give of some energy in the form of light. This light appears to be reflected from the surface of the metal which gives a shining look.

5. Malleable and ductile

Metals are malleable and ductile whenever stress is applied on them. Their layers slip pass each other. The structure of the metal changes without fracturing.

Important long questions from past papers

8. Define hydrogen bonding. Explain the structure of Ice on the basis of Hydrogen Boding.
9. Discuss hydrogen bonding in biological compounds.
10. Explain hydrogen bonding in NH_3, H_2O and HF.
11. Classify solids on the basis of Bonding.
12. Write brief not on isomorphism, anisotropy, polymorphism, transition temperature
13. What are ionic solids? Give their properties in detail.
14. What are covalent solids? Write their properties in detail.
15. What are liquid crystals? Give their uses in daily life.
16. What are molecular solids? Give their important characteristics?
17. Define vapor pressure? Discuss its measurement by Manometric method.
18. Discuss London dispersion forces. Elaborate factors on which it depends.
19. Give Postulates of Kinetic Molecular theory (K.M.T).