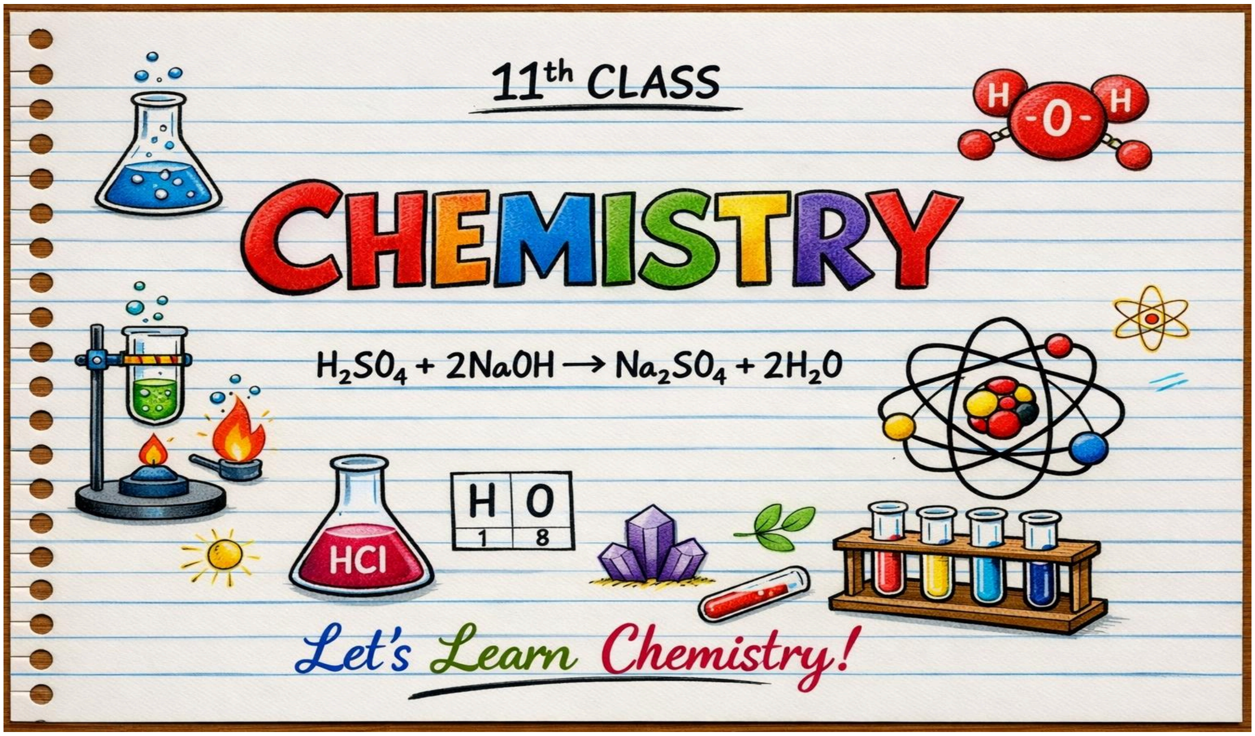


Class: 11th

Subject: Chemistry

Unit 5: STATES AND PHASES OF MATTER

❖ Important MCQs:

1. Matter exists in how many main states?

(a) Two

(b) Three

(c) Four ✅

(d) Five

2. Which state of matter is considered the simplest form?

(a) Solid

(b) Liquid

(c) Gas ✅

(d) Plasma

3. Most matter around us exists in which state?

(a) Liquid

(b) Solid ✅

(c) Gas

(d) Plasma

4. The liquid state of matter exists within a:

(a) Wide range of temperature and pressure

(b) Narrow range of temperature and pressure ✅

(c) Only high temperature range

(d) Only low pressure range

5. Gases have:

(a) Definite shape and volume

(b) Definite volume but no shape

(c) No definite shape and volume ✅

(d) Definite shape but no volume

6. The volume of a gas is equal to the:

(a) Volume of its particles

(b) Volume of the container ✅

(c) Twice the container volume

(d) Half of the container volume

7. In ideal gases, the volume of particles is considered:

(a) Large

(b) Small but significant

(c) Zero ✅

(d) Equal to container volume

8. Approximately how much empty space is present in gases?

(a) 50%

(b) 75%

(c) 99.9% ✅

(d) 100%

9. Gases are easily compressible because of:

(a) Strong intermolecular forces

(b) Small particle size

(c) Large empty space between particles ✅

(d) High density

10. Sudden expansion of a gas results in:

(a) Heating

(b) Cooling ✅

(c) No change

(d) Condensation

11. The cooling effect during expansion of gas is known as:

(a) Boyle's law effect

(b) Charles' law effect

(c) Joule-Thomson effect ✅

(d) Avogadro's law effect

12. Gas pressure is caused by:

(a) Attraction between molecules

(b) Repulsion between molecules

(c) Collisions of molecules with container walls ✅

(d) Molecular volume

13. In ideal gases, intermolecular forces are:

(a) Strong

(b) Moderate

(c) Negligible ✅

(d) Variable

14. According to Charles' law, volume is directly proportional to:

(a) Pressure

(b) Temperature (at constant pressure and moles) ✅

(c) Number of moles only

(d) Density

15. According to Avogadro's law, volume is directly proportional to:

(a) Pressure

(b) Temperature

(c) Number of moles (at constant T and P) ✅

(d) Mass

16. Which statement correctly describes diffusion in liquids?

(a) Faster than gases

(b) Slower than gases due to stronger intermolecular forces ✅

(c) Same as gases

(d) Not possible in liquids

17. The small spaces between liquid molecules are due to:

(a) Weak intermolecular forces

(b) Strong intermolecular forces keeping molecules closer together ✅

(c) High temperature only

(d) Absence of motion

18. Compared to gases, liquids are approximately:

(a) 10 times less compressible

(b) 100 times less compressible

(c) 10^5 times less compressible ✅

(d) 10^8 times less compressible

19. Molecules in a liquid exhibit:

(a) No motion

(b) Random motion with lower speed than gases ✅

(c) Only vibrational motion

(d) Fixed motion in lattice

20. Which observation proves that liquid molecules move freely?

(a) Crystallization

(b) Diffusion among miscible liquids ✅

(c) Freezing

(d) Sublimation

21. Expansion of liquids on heating is:

(a) Greater than gases

(b) Equal to solids

(c) Negligible compared to gases ✅

(d) Zero

22. According to kinetic molecular theory, increasing temperature of a liquid results in:

(a) Decrease in kinetic energy

(b) Increase in kinetic energy ✅

(c) No change in kinetic energy

(d) Conversion into solid directly

23. Conversion of a liquid into vapour occurs when:

(a) Kinetic energy decreases

(b) Intermolecular forces increase

(c) Kinetic energy increases sufficiently to overcome intermolecular forces ✅

(d) Volume decreases

24. Instantaneous dipole-induced dipole forces are also known as:

(a) Hydrogen bonding

(b) Dipole-dipole forces

(c) London dispersion forces ✅

(d) Ionic forces

25. Instantaneous dipole is formed due to:

(a) Permanent charge separation

(b) Temporary uneven distribution of electrons ✅

(c) Nuclear attraction

(d) Hydrogen bonding

26. London dispersion forces are present in:

(a) Polar molecules only

(b) Non-polar molecules only ✅

(c) Ionic compounds only

(d) All solids only

27. Strength of dispersion forces increases with:

(a) Decrease in molecular mass

(b) Increase in molecular mass and polarizability ✅

(c) Decrease in surface area

(d) Increase in temperature only

28. Which halogen has the strongest intermolecular forces?

(a) Fluorine (F_2)

(b) Chlorine (Cl_2)

(c) Bromine (Br_2)

(d) Iodine (I_2) ✅

29. Which factor increases intermolecular forces due to shape?

(a) Decrease in surface area

(b) Increase in surface area ✅

(c) Decrease in molecular mass

(d) Increase in temperature

30. Among isomers, which type shows highest boiling point?

(a) Highly branched molecules

(b) Molecules with smallest surface area

(c) Straight-chain molecules with largest surface area ✅

(d) Cyclic molecules only

31. Permanent dipole–dipole forces occur between:

(a) Non-polar molecules

(b) Polar molecules with permanent dipoles ✅

(c) Ionic compounds only

(d) Metals only

32. In HCl molecules, dipole–dipole forces arise due to:

(a) Equal sharing of electrons

(b) Attraction between positive and negative ends of polar molecules ✅

(c) Nuclear forces

(d) Hydrogen bonding only

33. Chloroform ($CHCl_3$) shows dipole–dipole forces because:

(a) It is non-polar

(b) It has a permanent dipole with partial charges at different ends ✅

(c) It has metallic bonding

(d) It has ionic bonding

34. Hydrogen bonding occurs when hydrogen is directly bonded to:

(a) Carbon, Sulphur, Phosphorus

(b) F, O, or N atoms ✅

(c) Chlorine only

(d) All non-metals

35. For hydrogen bonding, the electronegative atom must have:

(a) No electrons

(b) Lone pair of electrons ✅

(c) Only protons

(d) High mass

36. Hydrogen bonding is stronger than:

(a) Covalent, ionic, and metallic bonds

(b) Only ionic bonds

(c) Intermolecular forces but weaker than covalent bonds ✅

(d) Nuclear forces

37. Hydrogen bonds are represented by:

(a) Solid line

(b) Double line

(c) Dotted line (...) ✅

(d) Triple line

38. Water can form how many hydrogen bonds per molecule on average?

(a) One

(b) Two ✅

(c) Three

(d) Four

39. Ammonia (NH_3) forms fewer hydrogen bonds than water because:

(a) It has no hydrogen atoms

(b) It has only one lone pair on nitrogen ✅

(c) It is non-polar

(d) It has ionic bonding

40. The high boiling point of water is mainly due to:

(a) Weak intermolecular forces

(b) Strong hydrogen bonding between molecules ✅

(c) Ionic bonding

(d) Metallic bonding

41. Ice has lower density than water because:

(a) Molecules are tightly packed

(b) Hydrogen bonds break completely

(c) Empty spaces are created in the hydrogen-bonded structure ✅

(d) Molecules lose mass

42. When water freezes, its volume:

(a) Decreases

(b) Remains constant

(c) Increases by about 9% ✅

(d) Becomes zero

43. Ice floats on water because:

(a) It has higher mass

(b) It has lower density than liquid water ✅

(c) It has stronger ionic bonds

(d) It is heavier than water

44. High specific heat capacity of water is mainly due to:

(a) Weak intermolecular forces

(b) Strong hydrogen bonding between molecules ✅

(c) Ionic bonding

(d) Absence of motion

45. The specific heat capacity of water is approximately:

(a) 2.0 J/g°C

(b) 3.5 J/g°C

(c) 4.18 J/g°C ✅

(d) 10 J/g°C

46. High boiling point of water is due to:

(a) Weak London forces

(b) Strong hydrogen bonding requiring more energy to break ✅

(c) Low molecular mass

(d) Absence of intermolecular forces

47. Surface tension of a liquid is caused by:

(a) Outward force on molecules

(b) Inward attraction between surface molecules due to intermolecular forces ✅

(c) Gravitational force only

(d) Repulsive nuclear forces

48. Liquids tend to form spherical droplets because:

(a) They have high temperature

(b) They minimize surface area due to surface tension ✅

(c) They have high pressure

(d) They have no intermolecular forces

49. Surface tension generally decreases with increase in:

(a) Pressure

(b) Temperature ✅

(c) Density

(d) Volume

50. Viscosity of a liquid refers to:

(a) Its ability to evaporate

(b) Its resistance to flow ✅

(c) Its density

(d) Its compressibility

51. Which factor primarily determines the viscosity of a liquid?

(a) Volume of liquid

(b) Intermolecular forces ✅

(c) Color of liquid

(d) Density only

52. Why does viscosity decrease with increase in temperature?

(a) Molecules become larger

(b) Intermolecular forces become stronger

(c) Molecules gain kinetic energy and overcome intermolecular forces ✅

(d) Density of liquid increases

53. Which of the following liquids will have the highest viscosity?

(a) Weak intermolecular forces

(b) Medium intermolecular forces

(c) Strong intermolecular forces ✅

(d) No intermolecular forces

54. Evaporation is best described as:

(a) Boiling at fixed temperature

(b) Conversion of liquid to vapour at any temperature ✅

(c) Condensation of vapour

(d) Freezing of liquid

55. Why does evaporation cause cooling?

(a) Low energy molecules escape

(b) High energy molecules escape, lowering average kinetic energy of remaining molecules ✅

(c) Heat is released to surroundings

(d) Pressure increases in liquid

56. Which molecules escape first during evaporation?

(a) Low kinetic energy molecules

(b) High kinetic energy molecules ✅

(c) All molecules equally

(d) Heavy molecules only

57. Earthenware vessels keep water cool because:

(a) They absorb heat from water

(b) They prevent evaporation

(c) Water evaporates through pores causing cooling ✅

(d) They increase temperature of water

58. Vapour pressure is defined as:

(a) Pressure of liquid at boiling point

(b) Pressure exerted by solid above liquid

(c) Pressure exerted by vapour in equilibrium with its liquid at a given temperature ✅

(d) Atmospheric pressure above liquid

59. Vapour pressure of a liquid depends on:

(a) Surface area

(b) Amount of liquid

(c) Temperature and intermolecular forces ✅

(d) Shape of container

60. Which liquid will have the highest vapour pressure at the same temperature?

(a) Strong intermolecular forces

(b) Weak intermolecular forces ✅

(c) High viscosity

(d) High density

61. Increase in temperature causes vapour pressure to:

(a) Decrease

(b) Remain constant

(c) Increase ✅

(d) Become zero

62. In a closed container, evaporation stops when:

(a) Liquid is finished

(b) Temperature reaches zero

(c) Rate of evaporation equals rate of condensation ✅

(d) Pressure becomes zero

63. Condensation is the process in which:

(a) Liquid changes into vapour

(b) Vapour changes into liquid ✅

(c) Solid changes into liquid

(d) Liquid changes into solid

64. Which statement is correct regarding surface area and vapour pressure?

(a) Vapour pressure depends on surface area

(b) Vapour pressure increases with surface area

(c) Vapour pressure does not depend on surface area ✅

(d) Vapour pressure decreases with surface area

65. Liquids with stronger intermolecular forces show:

(a) Higher vapour pressure and higher evaporation

(b) Lower vapour pressure and lower evaporation rate ✅

(c) Higher evaporation rate only

(d) No effect on vapour pressure

66. Boiling point is the temperature at which:

(a) Vapour pressure is zero

(b) Vapour pressure equals external pressure ✅

(c) Intermolecular forces become zero

(d) Temperature becomes maximum permanently

67. During boiling, bubbles formed inside the liquid rise because:

(a) Their density is higher than liquid

(b) Internal vapour pressure is greater than external pressure ✅

(c) External pressure pulls them upward

(d) Temperature decreases suddenly

68. A constant stream of bubbles at boiling point indicates that:

(a) Evaporation has stopped

(b) Vapour pressure is less than external pressure

(c) Vapour pressure equals external pressure and equilibrium is maintained dynamically ✅

(d) Liquid has become solid

69. At boiling point, further heating does not increase temperature because:

(a) Heat is lost completely

(b) Heat is used to increase pressure

(c) Heat is used to overcome intermolecular forces ✅

(d) Molecules stop moving

70. Which energy change occurs during boiling after reaching boiling point?

(a) Increase in kinetic energy only

(b) Decrease in kinetic energy

(c) Energy is used to break intermolecular forces ✅

(d) No energy change occurs

71. Liquids with stronger intermolecular forces generally have:

(a) Higher vapour pressure and lower boiling point

(b) Lower vapour pressure and higher boiling point ✅

(c) Lower vapour pressure and lower boiling point

(d) Higher vapour pressure and higher boiling point

72. Why does water have a higher boiling point compared to ethanol?

(a) Weaker intermolecular forces

(b) Stronger intermolecular forces in water ✅

(c) Lower molecular mass

(d) Higher density only

73. Increase in external pressure causes boiling point to:

(a) Decrease

(b) Remain unchanged

(c) Increase ✅

(d) Become zero

74. Water boils at 120°C at 1489 torr because:

(a) Intermolecular forces decrease

(b) External pressure is higher than normal atmospheric pressure ✅

(c) Temperature of surroundings increases

(d) Vapour pressure decreases to zero

75. At Mount Everest, water boils at a lower temperature because:

(a) High external pressure

(b) Low external pressure ✅

(c) High humidity

(d) High kinetic energy of air molecules

76. The function of a pressure cooker is to:

(a) Decrease temperature of boiling

(b) Increase external pressure to raise boiling point ✅

(c) Reduce intermolecular forces

(d) Allow vapour to escape freely

77. Food cooks faster in a pressure cooker because:

(a) Boiling point decreases

(b) Boiling point increases due to increased pressure, allowing higher temperature cooking ✅

(c) Heat is lost rapidly

(d) Evaporation stops completely

78. Energy change during a physical change is mainly due to:

(a) Change in mass

(b) Change in intermolecular forces ✅

(c) Change in volume only

(d) Change in color

79. Molar heat of fusion is defined as:

(a) Heat required to vaporize one mole of liquid

(b) Heat required to melt one mole of solid at its melting point at 1 atm ✅

(c) Heat required to cool one mole of liquid

(d) Heat released during freezing

80. The molar heat of fusion of water is approximately:

(a) 40.6 kJ/mol

(b) 4.6 kJ/mol ✅

(c) 21.7 kJ/mol

(d) 15.6 kJ/mol

81. High heat of fusion of water contributes to:

(a) Rapid evaporation of ice

(b) Stability of glaciers and ice caps ✅

(c) Increase in vapor pressure

(d) Decrease in melting point

82. If the heat of fusion of water were low, then:

(a) Ice would be more stable

(b) Glaciers would melt easily causing environmental changes ✅

(c) Boiling point would increase

(d) Vapour pressure would decrease

83. Molar heat of vaporization is defined as:

(a) Heat required to melt one mole of solid

(b) Heat required to convert one mole of liquid into vapour at its boiling point at 1 atm ✅

(c) Heat released during condensation

(d) Heat required to freeze one mole of liquid

84. The molar heat of vaporization of water is approximately:

(a) 4.6 kJ/mol

(b) 15.6 kJ/mol

(c) 40.6 kJ/mol ✅

(d) 21.7 kJ/mol

85. Lower molar heat of vaporization in NH_3 and HCl compared to water is due to:

(a) Weaker intermolecular forces in these substances ✅

(b) Higher molecular mass of water

(c) Greater density of NH_3

(d) Lower temperature of water

86. During phase change from liquid to gas, particles:

(a) Lose energy and come closer

(b) Gain energy and move farther apart ✅

(c) Stop moving

(d) Become fixed in position

87. Which property is NOT significant in solids?

(a) Definite shape

(b) Definite volume

(c) High compressibility ✅

(d) Strong intermolecular forces

88. Solids are difficult to compress because:

(a) Particles are far apart

(b) Particles are closely packed with negligible space between them ✅

(c) Particles are in random motion

(d) Forces between particles are weak

89. Particles in solids mainly possess:

(a) Translational motion

(b) Rotational motion

(c) Vibrational motion about fixed positions ✅

(d) No motion at all

90. Which type of forces hold particles together in solids?

(a) Only gravitational forces

(b) Ionic, covalent, metallic, and van der Waals forces ✅

(c) Only magnetic forces

(d) Nuclear forces only

91. A substance shows a definite geometrical shape and sharp melting point. It is most likely:

(a) Amorphous solid

(b) Crystalline solid ✅

(c) Liquid crystal

(d) Gas

92. Which property confirms long-range order in a solid?

(a) Irregular melting

(b) Random particle arrangement

(c) Constant interfacial angles in all crystals of the same substance ✅

(d) High viscosity

93. Why do crystalline solids show definite interfacial angles?

(a) Particles are loosely packed

(b) Particles are arranged in a regular repeating 3D pattern ✅

(c) Particles are in random motion

(d) Due to high temperature

94. A solid breaks along specific planes. This indicates the presence of:

(a) Amorphous structure

(b) Cleavage planes in crystalline solids ✅

(c) Liquid behavior

(d) Gas expansion

95. Which condition is essential for crystal growth?

(a) Dilute solution

(b) Supersaturated or saturated solution with slow cooling/evaporation ✅

(c) High pressure gas

(d) Low temperature solid only

96. Addition of a seed crystal to a supersaturated solution triggers crystallization because:

(a) It dissolves the solute

(b) It disturbs the metastable equilibrium and provides a nucleation site ✅

(c) It lowers temperature instantly

(d) It increases pressure

97. Which statement best explains why amorphous solids lack sharp melting points?

(a) Uniform particle arrangement

(b) Presence of long-range order

(c) Irregular and non-uniform particle arrangement requiring a range of energy to break interactions ✅

(d) High vapor pressure

98. Crystallites in amorphous solids are:

(a) Large perfect crystals

(b) Small regions with short-range order within an otherwise disordered structure ✅

(c) Liquid droplets

(d) Gas pockets

99. Glass is classified as an amorphous solid because:

(a) It has sharp melting point

(b) It has long-range periodic structure

(c) It lacks long-range order and shows gradual softening on heating ✅

(d) It is highly crystalline

100. Amorphous solids can be molded into different shapes because:

(a) They have strong rigid lattice

(b) They lack long-range order and can flow under heat/pressure over time ✅

(c) They are liquids

(d) They have high melting point

101. Liquid crystals are best described as substances that:

(a) Are fully solid

(b) Are fully liquid

(c) Exhibit properties intermediate between solids and liquids ✅

(d) Are gases under pressure

102. The molecular shape most commonly associated with liquid crystals is:

(a) Spherical

(b) Cubical

(c) Rod-like elongated molecules ✅

(d) Irregular clusters

103. Why do liquid crystal molecules align parallel to each other?

(a) Due to gravitational force

(b) Due to their elongated shape and intermolecular interactions favoring ordered orientation ✅

(c) Due to external pressure only

(d) Due to random motion

104. In liquid crystals, molecules can:

(a) Only vibrate

(b) Only rotate end-over-end

(c) Move like liquids but have restricted rotational motion about their long axis ✅

(d) Remain completely fixed

105. Which property of liquid crystals makes them useful in LCDs?

(a) Electrical conductivity

(b) Ability to modulate light under electric fields due to optical anisotropy ✅

(c) High density

(d) High melting point

106. Liquid crystals show optical properties because:

(a) They are fully amorphous

(b) Their molecular arrangement interacts with light differently depending on orientation ✅

(c) They have no molecular structure

(d) They are opaque

107. Liquid crystals are used in temperature sensing because:

(a) Their color changes with temperature due to structural changes affecting light reflection ✅

(b) They evaporate quickly

(c) They become gases

(d) They conduct electricity highly

108. In medical diagnostics, liquid crystals help detect tumors because:

(a) Tumors are colder than surrounding tissue

(b) Tumors often produce localized temperature differences detectable by liquid crystals ✅

(c) Tumors change mass

(d) Tumors change density only

109. The working principle of LCD screens is based on:

(a) Chemical reactions

(b) Light modulation by controlled orientation of liquid crystal molecules under electric fields ✅

(c) Heat conduction

(d) Sound waves

110. Which of the following is a correct comparison between crystalline and amorphous solids?

(a) Crystalline: random arrangement; Amorphous: ordered arrangement

(b) Crystalline: sharp melting point; Amorphous: melts over a range ✅

(c) Crystalline: no structure; Amorphous: definite structure

(d) Both have identical properties

❖ Important Short Questions:

1. Define intermolecular forces. Explain their types with examples.

Ans: Intermolecular forces are the attractive forces that exist between molecules of a substance. These forces are weaker than chemical bonds but determine physical properties like boiling point and viscosity.

Types of intermolecular forces:

1. London dispersion forces: Present in all molecules, especially non-polar molecules.

Example: Forces between helium atoms, nitrogen (N_2), methane (CH_4).

2. Dipole-dipole forces: Occur between polar molecules having permanent dipoles.

Example: Interaction between HCl molecules.

3. Hydrogen bonding: Strong intermolecular force occurring when hydrogen is bonded with highly electronegative atoms like N, O, or F.

Example: Water (H_2O), ammonia (NH_3).

2. Differentiate between intermolecular and intramolecular forces.

Answer:

Intermolecular forces:

- Forces between molecules
- Weaker in nature
- Control physical properties (boiling point, melting point)
- **Example:** Forces between water molecules

Intramolecular forces:

- Forces within a molecule (chemical bonds)
- Very strong
- Hold atoms together in a molecule
- **Example:** O–H bonds within a water molecule

3. What are London dispersion forces? Explain with examples.

Answer:

London dispersion forces are weak attractive forces caused by temporary dipoles formed due to movement of electrons.

They occur in all atoms and molecules, especially non-polar ones.

Examples:

- Noble gases like helium (He), neon (Ne)
- Non-polar molecules like oxygen (O_2), nitrogen (N_2), methane (CH_4)

Even though these molecules are non-polar, temporary dipoles create weak attractions between them.

4. Explain dipole-dipole interactions with suitable examples.

Answer:

Dipole-dipole interactions occur between molecules that have permanent dipoles (polar molecules). The positive end of one molecule attracts the negative end of another.

Examples:

- Hydrogen chloride (HCl) molecules attract each other
- Sulfur dioxide (SO_2) molecules

These forces are stronger than London dispersion forces but weaker than hydrogen bonding.

5. What is hydrogen bonding? Why is it stronger than other intermolecular forces?

Answer:

Hydrogen bonding is a strong intermolecular force that occurs when hydrogen is bonded to highly electronegative atoms like nitrogen (N), oxygen (O), or fluorine (F).

Examples:

- Water (H_2O)
- Ammonia (NH_3)
- Hydrogen fluoride (HF)

Reason it is stronger:

- Large electronegativity difference between H and N/O/F
- Creates highly polar bonds
- Leads to strong attraction between molecules

Therefore, hydrogen bonding is stronger than dipole-dipole and London forces.

6. How do intermolecular forces affect physical properties of substances?

Answer:

Intermolecular forces significantly influence physical properties:

1. Boiling point: Stronger forces → higher boiling point

Example: Water has a higher boiling point than methane due to hydrogen bonding.

2. Melting point: Stronger forces require more energy to break

Example: Ionic and hydrogen-bonded substances have higher melting points.

3. Viscosity: Stronger forces → higher viscosity

Example: Honey is more viscous than water.

Vapour pressure: Stronger forces → lower vapour pressure

Example: Water has lower vapour pressure than ether.

Surface tension: Stronger forces → higher surface tension

Example: Water forms droplets due to strong cohesion between molecules.

7. Why does a liquid not increase in temperature at its boiling point during continuous heating?

Answer:

At the boiling point, the heat supplied is used to overcome intermolecular forces rather than increasing the kinetic energy of molecules.

This energy is called latent heat of vaporization.

Explanation:

- Temperature depends on kinetic energy.
- During boiling, added heat breaks intermolecular attractions instead of increasing motion.
- **Therefore,** temperature remains constant until the entire liquid changes into vapour.

Example: Water at 100°C keeps boiling at the same temperature until it completely vaporizes.

8. Explain the role of intermolecular forces in determining boiling point.

Answer:

Boiling point depends on the strength of intermolecular forces:

- Stronger intermolecular forces ⟶ higher boiling point
- Weaker intermolecular forces ⟶ lower boiling point

Explanation: More energy is required to separate molecules when forces are strong.

Examples:

- Water (H_2O) has hydrogen bonding ⟶ high boiling point (100°C)
- Methane (CH_4) has weak dispersion forces ⟶ very low boiling point (−161°C)

9. How does external pressure affect the boiling point of a liquid? Explain with examples.

Answer:

Boiling occurs when vapour pressure equals external pressure.

- Increase in external pressure ⟶ increase in boiling point
- Decrease in external pressure ⟶ decrease in boiling point

Explanation: Higher external pressure requires higher vapour pressure (thus higher temperature) to boil.

Examples:

- In a pressure cooker, boiling point increases ⟶ food cooks faster
- At high altitudes (low pressure), water boils at lower temperature

10. Why does water boil at lower temperature on mountains compared to sea level?

Answer:

At high altitudes (mountains):

- Atmospheric pressure is lower
- Therefore, water needs less temperature to reach boiling point

Explanation: Boiling occurs when vapour pressure equals external pressure. Since external pressure is low, this condition is achieved at a lower temperature.

Example:

- Water may boil around 90–95°C in mountainous areas instead of 100°C.

11. Describe the working of a pressure cooker in terms of boiling point.

Answer:

A pressure cooker works by increasing the internal pressure inside the vessel.

Explanation:

- Steam is trapped inside, increasing pressure
- Higher pressure raises the boiling point of water
- Water can reach temperatures above 100°C
- This higher temperature cooks food faster

Example:

Rice and meat cook more quickly in a pressure cooker due to higher temperature inside.

12. Define evaporation. Explain factors affecting evaporation.

Answer:

Evaporation is the process in which liquid molecules escape from the surface and change into vapour at temperatures below the boiling point.

Factors affecting evaporation:

1. Temperature: Higher temperature ⟶ faster evaporation

Example: Water evaporates faster in summer.

2. Surface area: Larger surface area ⟶ faster evaporation

Example: Water in a plate evaporates faster than in a glass.

3. Humidity: Higher humidity ⟶ slower evaporation

Example: Clothes dry slowly in humid weather.

4. Wind speed: Higher wind speed ⟶ faster evaporation

Example: Clothes dry faster on a windy day.

13. Why does evaporation cause cooling? Explain with kinetic theory.

Answer:

Evaporation causes cooling because high-energy molecules escape from the liquid surface.

Kinetic theory explanation:

- Molecules have different kinetic energies
- Fast-moving (high-energy) molecules escape first
- Remaining molecules have lower average kinetic energy
- Since temperature depends on average kinetic energy, it decreases

Example: Sweating cools the body as sweat evaporates from the skin.

14. Define vapour pressure. What factors affect it?

Answer:

Vapour pressure is the pressure exerted by vapour in equilibrium with its liquid in a closed container at a given temperature.

Factors affecting vapour pressure:

- **Nature of liquid (intermolecular forces):** Strong forces ⟶ lower vapour pressure
- **Temperature:** Higher temperature ⟶ higher vapour pressure
- **Volatility of liquid:** More volatile liquids have higher vapour pressure

Example: Ether has higher vapour pressure than water.

15. How does temperature affect vapour pressure?

Answer:

As temperature increases:

- Kinetic energy of molecules increases
- More molecules escape into vapour phase
- Vapour pressure increases

Example: Vapour pressure of water at 80°C is higher than at 25°C.

16. Why do volatile liquids have higher vapour pressure?

Answer:

Volatile liquids have weak intermolecular forces, so molecules can easily escape into vapour.

Explanation:

- Weak forces require less energy to overcome
- More molecules enter vapour phase
- This increases vapour pressure

Examples: Ether, alcohol → high vapour pressure; water → comparatively lower vapour pressure

Conclusion:

Weaker intermolecular forces → higher volatility → higher vapour pressure.

17. Define boiling point and explain its relation with vapour pressure.

Answer:

Boiling point is the temperature at which a liquid starts boiling, i.e., when its vapour pressure becomes equal to the external (atmospheric) pressure.

Relation with vapour pressure:

- A liquid boils when its vapour pressure equals external pressure

- At this point, bubbles of vapour form throughout the liquid

Example: Water boils at 100°C at 1 atm because its vapour pressure equals atmospheric pressure at that temperature.

18. Why does boiling occur when vapour pressure equals external pressure?

Answer:

Boiling occurs when:

- Vapour pressure inside the liquid becomes equal to external pressure
- At this point, vapour bubbles can form and rise without collapsing

Explanation:

- If external pressure is higher, bubbles collapse
- When vapour pressure matches external pressure, bubbles remain stable and escape as vapour

19. How does external pressure affect boiling point?

Answer:

- Increase in external pressure → increase in boiling point
- Decrease in external pressure → decrease in boiling point

Explanation: Higher external pressure requires higher temperature for vapour pressure to match it.

Example: In a pressure cooker, increased pressure raises the boiling point of water.

20. Explain why boiling point decreases at high altitudes.

Answer:

At high altitudes:

- Atmospheric pressure is lower
- Therefore, vapour pressure equals external pressure at a lower temperature

Explanation: Less external pressure means less heat is needed for boiling.

Example: Water may boil below 100°C on mountains.

21. Describe the working of a pressure cooker in terms of boiling point.

Answer:

A pressure cooker works by increasing the pressure inside the container.

Explanation:

- Steam is trapped, increasing internal pressure
- Higher pressure raises the boiling point of water above 100°C
- Water reaches higher temperature before boiling

- This cooks food faster

Example: Food cooks more quickly in a pressure cooker due to higher boiling temperature.

22. Define molar heat of fusion and molar heat of vaporization.

Answer:

Molar heat of fusion:

The amount of heat required to convert 1 mole of a solid into liquid at its melting point without temperature change.

Molar heat of vaporization:

The amount of heat required to convert 1 mole of a liquid into vapour at its boiling point without temperature change.

Example: Ice ⟶ water (fusion), water ⟶ steam (vaporization).

23. Why is heat of vaporization greater than heat of fusion?

Answer:

- In melting, molecules only loosen their arrangement.
- In vaporization, molecules must completely separate and escape into gas phase.
- Stronger intermolecular forces must be fully overcome in vaporization.

Conclusion: More energy is required for vaporization than fusion.

24. Explain energy changes during melting and boiling.

Answer:

- **During melting,** heat energy is absorbed to weaken intermolecular forces.
- **During boiling,** heat energy is used to completely overcome intermolecular forces and convert liquid into gas.

In both cases, energy is used for breaking intermolecular attractions, not increasing temperature.

25. Why does temperature remain constant during phase change?

Answer:

- Heat supplied is used as latent heat to break intermolecular forces.
- No increase in kinetic energy of particles occurs.
- Since temperature depends on kinetic energy, it remains constant.

Example: Water stays at 100°C during boiling until all liquid becomes vapor.

26. Define crystalline solids and describe their properties.

Answer:

Crystalline solids have a regular, orderly, repeating arrangement of particles.

Properties:

- Definite geometric shape
- Sharp melting point
- Long-range order
- Anisotropic nature
- Cleave along definite planes

Examples: NaCl, sugar, quartz

27. What is a crystal lattice? Explain its importance.

Answer:

A crystal lattice is a 3D regular arrangement of particles in a crystalline solid.

Importance:

- Determines structure and shape
- Controls physical properties like melting point and density
- Explains orderly arrangement of particles

28. What are amorphous solids? Give examples.

Answer:

Amorphous solids have irregular and random arrangement of particles.

Examples: Glass, rubber, plastics

Properties:

- No definite shape
- No sharp melting point (soften over a range)
- Isotropic nature

29. Differentiate between crystalline and amorphous solids (concept-based).

Answer:

Arrangement:

- Crystalline → orderly; Amorphous → irregular

Melting point:

- Crystalline → sharp; Amorphous → gradual

Nature:

- Crystalline → anisotropic; Amorphous → isotropic

Examples:

Crystalline → NaCl; Amorphous → glass

30. Explain cleavage planes in crystalline solids.

Answer:

Cleavage planes are specific planes along which crystalline solids break easily due to their regular structure.

Example:

Salt crystals break into regular geometric shapes along cleavage planes.

31. What are isotropic and anisotropic solids? Explain with examples.

Answer:

Isotropic solids:

- Properties are the same in all directions
- Example: Amorphous solids like glass

Anisotropic solids:

- Properties vary with direction
- Example: Crystalline solids like quartz

32. Explain why water has higher boiling point than methane.

Answer:

- Water has strong hydrogen bonding
- Methane has weak London dispersion forces
- Stronger intermolecular forces require more energy to break

Conclusion: Water has a higher boiling point due to stronger intermolecular forces.

33. Why does ice float on water? Explain scientifically.

Answer:

- Ice has an open hexagonal structure due to hydrogen bonding

- This structure increases volume and decreases density
- Density of ice is less than liquid water

Conclusion: Ice floats because it is less dense than water.

34. Compare solids, liquids, and gases in terms of intermolecular forces and particle motion.

Answer:

Solids:

- Strong intermolecular forces, particles vibrate in fixed positions

Liquids:

- Moderate forces, particles can slide past each other

Gases:

- Very weak forces, particles move freely at high speed

35. Explain the relationship between intermolecular forces and physical state of matter.

Answer:

- Strong intermolecular forces ⟶ solid state
- Moderate forces ⟶ liquid state
- Weak forces ⟶ gaseous state

Conclusion: The strength of intermolecular forces determines whether a substance exists as solid, liquid, or gas.

❖ Important Long Questions

🌟 **Q1. Explain intermolecular forces in detail. Discuss their types and their effects on physical properties of substances.**

❖ Introduction

Intermolecular forces are the forces of attraction between molecules of a substance. These forces are much weaker than chemical (intramolecular) bonds but play a very important role in determining the physical properties of substances such as boiling point, melting point, viscosity, and surface tension.

✦ Types of Intermolecular Forces

1. London Dispersion Forces (id–id)

These are the weakest intermolecular forces and are present in all molecules, especially non-polar ones.

Explanation:

- Due to continuous movement of electrons, temporary dipoles are formed
- These temporary dipoles induce dipoles in nearby molecules
- This results in weak attraction

Examples:

- Helium (He), Neon (Ne)

- Non-polar molecules like O_2, N_2, CH_4

2. Dipole–Dipole Forces (pd–pd)

These forces exist between polar molecules having permanent dipoles.

Explanation:

- Positive end of one molecule attracts the negative end of another
- Stronger than London forces

Examples:

- Hydrogen chloride (HCl)
- Sulfur dioxide (SO_2)

3. Hydrogen Bonding

This is the strongest type of intermolecular force.

Conditions:

- Hydrogen must be bonded to highly electronegative atoms (F, O, N)
- The electronegative atom must have lone pair of electrons

Explanation:

- Strong polarity develops
- Hydrogen forms a strong attraction with nearby molecules

Examples:

- Water (H_2O)

- Ammonia (NH_3)
- Hydrogen fluoride (HF)

✦ Effects of Intermolecular Forces on Physical Properties

1. Boiling Point

- Stronger intermolecular forces ⟶ higher boiling point
- **Example:** Water has higher boiling point than methane

2. Melting Point

- Strong forces require more energy to break
- **Example:** Ice melts at a higher temperature than substances with weak forces

3. Viscosity

- Stronger forces ⟶ higher resistance to flow
- **Example:** Honey is more viscous than water

4. Vapour Pressure

- Stronger forces ⟶ lower vapour pressure
- **Example:** Water has lower vapour pressure than ether

5. Surface Tension

- Strong forces ⟶ higher surface tension
- Example: Water forms droplets due to strong cohesion

Conclusion

Intermolecular forces determine the state of matter and physical properties of substances. Stronger forces lead to solids and liquids, while weaker forces result in gases.

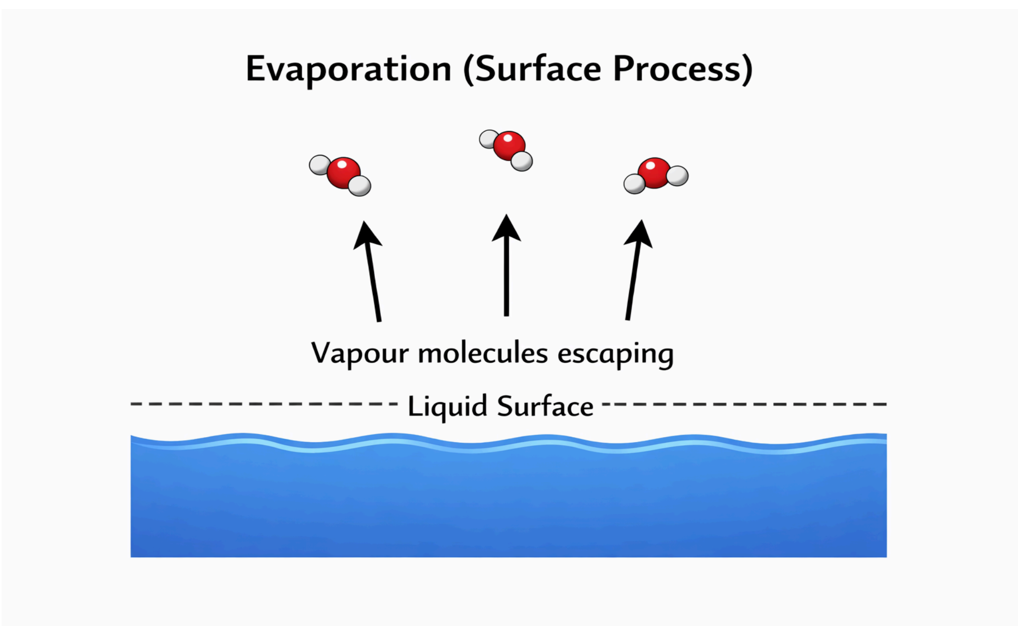
🌟 **Q2. Define evaporation, vapour pressure, and boiling point. Explain their relationship and factors affecting each with suitable examples.**

❖ **Answer:**

Introduction

Evaporation, vapour pressure, and boiling point are closely related concepts that describe the conversion of liquid into vapour and depend on temperature and intermolecular forces.

1. Evaporation



Definition:

Evaporation is the process in which liquid molecules escape from the surface and change into vapour at any temperature below boiling point.

Explanation:

- Only high-energy molecules escape from the surface
- Occurs slowly and at the surface only

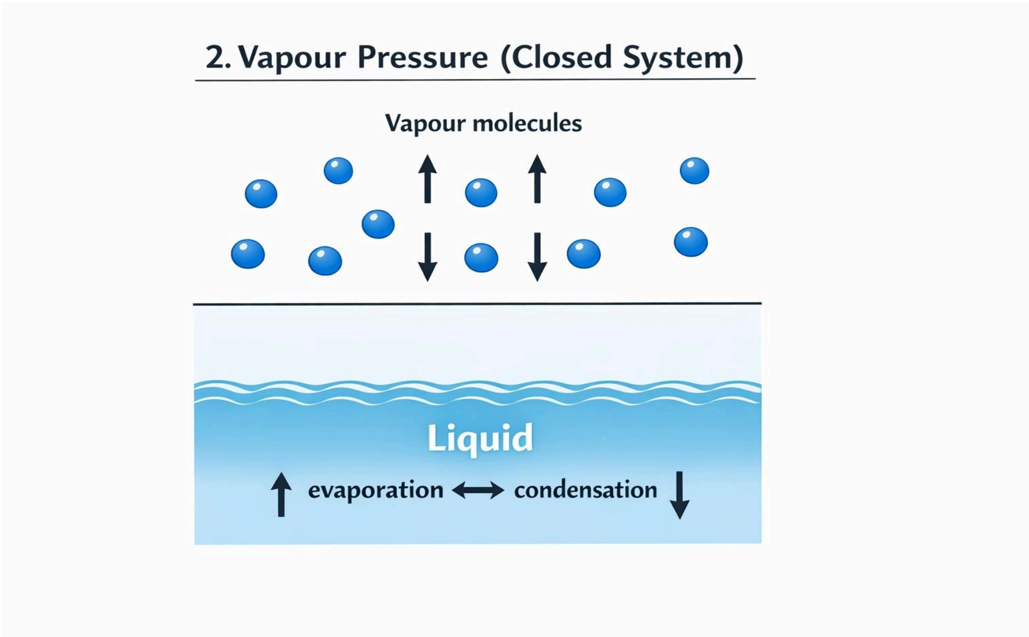
Examples:

- Drying of clothes
- Cooling of water in earthen pots

Factors affecting evaporation:

- **Temperature:** Higher temperature ⟶ faster evaporation
- **Surface area:** Larger area ⟶ faster evaporation
- **Humidity:** Higher humidity ⟶ slower evaporation
- **Wind speed:** Higher wind ⟶ faster evaporation

2. Vapour Pressure



Definition:

Vapour pressure is the pressure exerted by vapour in equilibrium with its liquid in a closed container at a given temperature.

Explanation:

- In a closed system, evaporation and condensation occur simultaneously
- At equilibrium, both rates become equal

Examples:

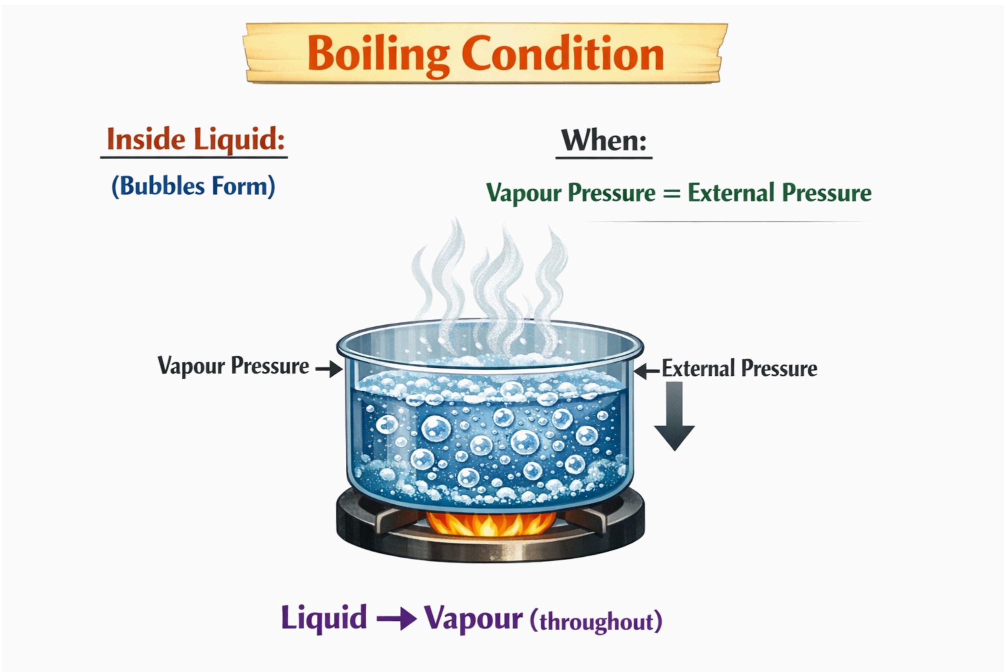
- Water in a closed bottle produces vapour pressure
- Ether shows higher vapour pressure than water

Factors affecting vapour pressure:

- **Temperature:** Increase → increases vapour pressure

- **Intermolecular forces:** Strong forces ⟶ lower vapour pressure
- **Nature of liquid (volatility):** More volatile ⟶ higher vapour pressure

3. Boiling Point



Definition:

Boiling point is the temperature at which vapour pressure of a liquid becomes equal to external pressure.

Explanation:

- At this point, bubbles form throughout the liquid
- Liquid changes into vapour rapidly

Examples:

- Water boils at 100°C at 1 atm
- Boiling point decreases on mountains

Factors affecting boiling point:

- **Intermolecular forces:** Stronger forces ⟶ higher boiling point
- **External pressure:** Higher pressure ⟶ higher boiling point

Relationship Between Evaporation, Vapour Pressure, and Boiling Point

- Evaporation increases vapour in air
- This leads to vapour pressure formation
- As temperature increases ⟶ vapour pressure increases
- When vapour pressure = external pressure ⟶ boiling occurs

Simple Flow: Evaporation ⟶ Vapour formation ⟶ Vapour pressure ⬆️ ⟶ Boiling point reached

Conclusion:

These three concepts are interconnected. Evaporation produces vapour, vapour pressure builds up, and boiling occurs when vapour pressure equals external pressure.

🌟 Q3. Describe viscosity and surface tension of liquids. Explain the factors affecting them and their importance in daily life.

❖ **Answer:**

Introduction

Viscosity and surface tension are important properties of liquids that arise due to intermolecular forces. They explain how liquids flow and how their surface behaves.

1. Viscosity

Definition:

Viscosity is the resistance of a liquid to flow.

Explanation:

- Molecules attract each other
- During flow, they resist movement
- Strong attraction ⟶ difficult flow ⟶ high viscosity

Factors Affecting Viscosity

1. Intermolecular Forces

- Strong forces ⟶ molecules tightly held
- Movement becomes difficult

👉 **Result:** Higher viscosity

2. Temperature

- Increases kinetic energy
- Molecules move freely

👉 **Result:** Viscosity decreases

3. Molecular Size

- Larger molecules create more resistance

👉 **Result:** Higher viscosity

Importance (Daily Life)

- Lubricants in machines
- Blood circulation
- Flow of oils and fuels
- Food thickness (honey, syrup)

2. Surface Tension

Definition:

Surface tension is the inward force acting on the surface of a liquid, making it behave like a stretched sheet.

Explanation:

- Surface molecules are pulled inward
- This creates a tight surface layer

Factors Affecting Surface Tension

1. Intermolecular Forces

- Strong forces ⟶ strong inward pull

👉 **Result:** High surface tension

2. Temperature

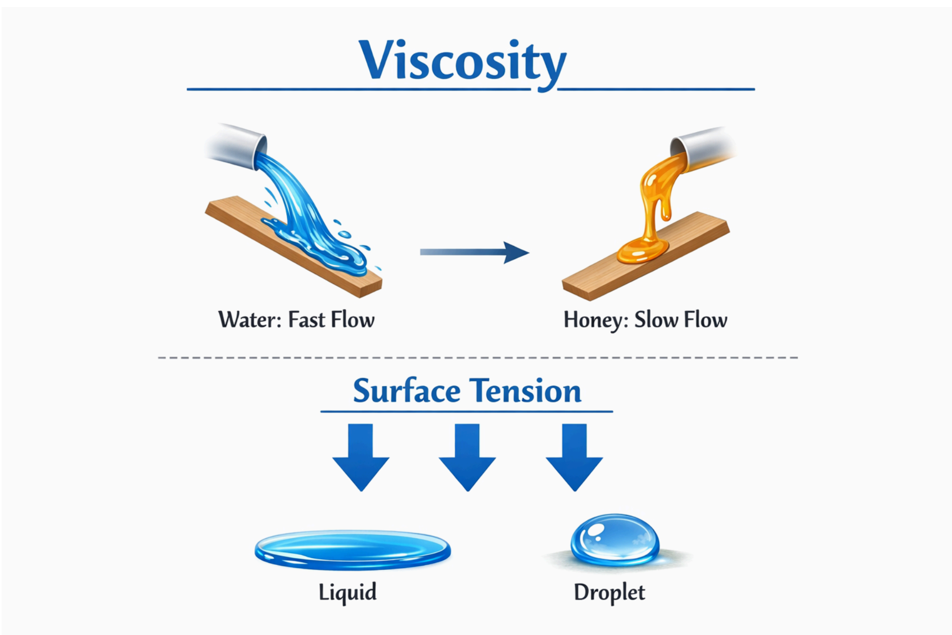
- Weakens intermolecular forces

👉 **Result:** Surface tension decreases

Importance (Daily Life)

- Cleaning by soaps
- Capillary action in plants
- Formation of droplets
- Ink and paint spreading

📊 **Diagram:**



✦ Summary:

- **Viscosity** = resistance to flow
- **Surface tension** = inward pull on liquid surface
- Both depend on intermolecular forces

- Strong forces → high viscosity & high surface tension
- Temperature ⬆️ → both decrease
- Both properties are important in daily life and industry

🌟 Q4. Explain molar heat of fusion and molar heat of vaporization. Discuss energy changes during phase changes and their effects on matter.

❖ **Answer:**

Introduction

When a substance changes its state (solid → liquid → gas), energy is either absorbed or released. This energy is related to the strength of intermolecular forces. The quantities used to measure this energy are called molar heat of fusion and molar heat of vaporization.

1. Molar Heat of Fusion (ΔHf)

Definition:

Molar heat of fusion is the amount of heat required to convert 1 mole of a solid into liquid at its melting point without any change in temperature.

Explanation:

- In solid state, particles are tightly packed
- When heat is supplied, particles gain energy

- This energy is used to overcome intermolecular forces, not to increase temperature
- As a result, solid changes into liquid

Example:

Ice ⟶ Water at 0°C

(Heat absorbed but temperature remains constant)

2. Molar Heat of Vaporization (ΔHv)

Definition:

Molar heat of vaporization is the amount of heat required to convert 1 mole of a liquid into vapour at its boiling point without temperature change.

Explanation:

- In liquid, molecules are loosely held
- To convert into gas, molecules must completely separate
- Large amount of energy is required to overcome intermolecular forces

Example:

Water ⟶ Steam at 100°C

3. Why ΔHv > ΔHf (Important Concept)

- In melting ⟶ molecules only loosen their positions

- In vaporization → molecules completely escape into gas phase

👉 **Result:**

More energy is required in vaporization than fusion

4. Energy Changes During Phase Changes

During Melting (Solid → Liquid):

- Heat is absorbed
- Intermolecular forces weaken
- Particles move more freely

During Boiling (Liquid → Gas):

- Large amount of heat is absorbed
- Intermolecular forces are completely broken
- Particles move freely at high speed

Important Point:

During both processes:

- Temperature remains constant
- Energy is used as latent heat (to break forces)

5. Effects of Energy Changes on Matter

1. Change in State

- Solid → Liquid → Gas

2. Increase in Particle Movement

- Particles gain freedom of motion

3. Increase in Distance Between Particles

- Particles move far apart in gas phase

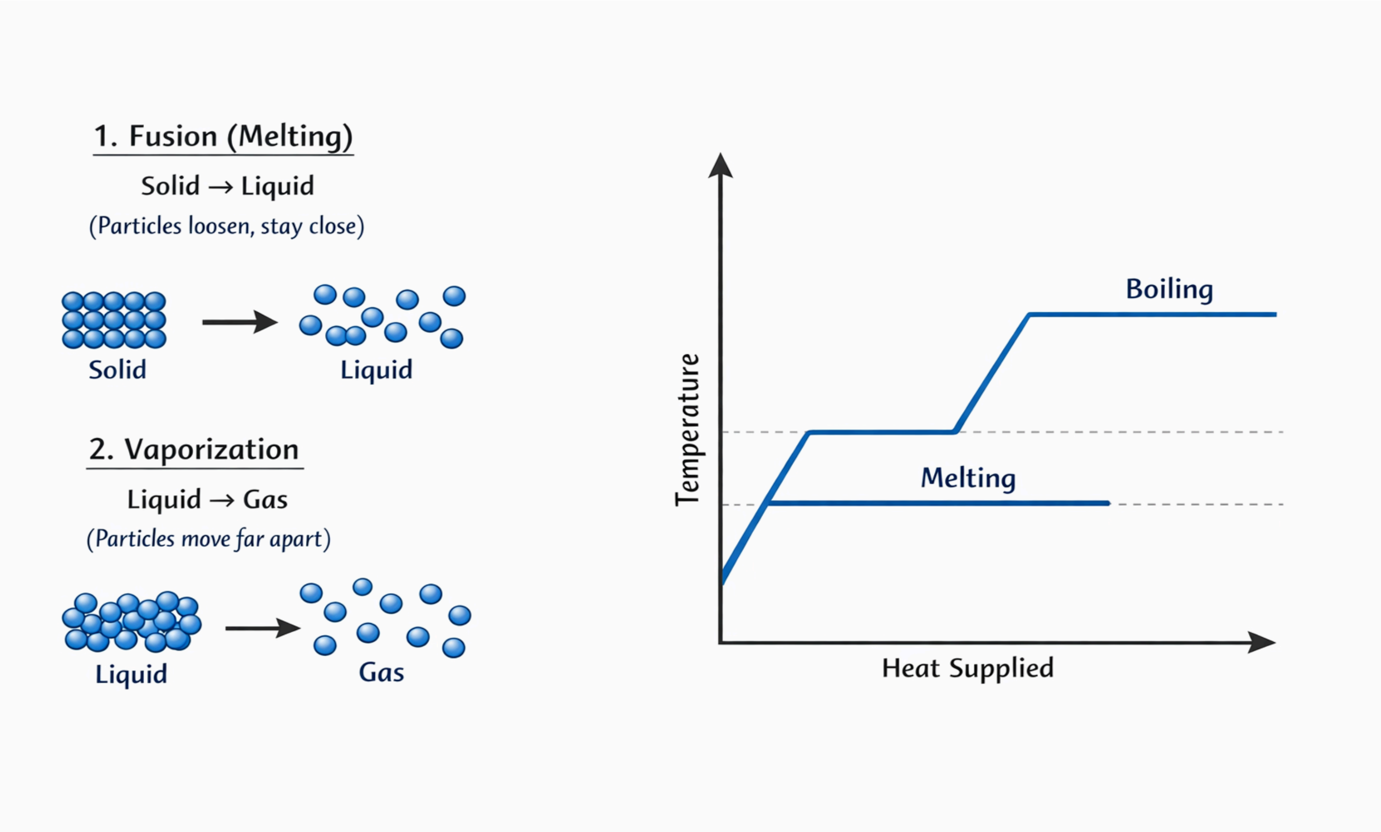
4. Change in Physical Properties

- Density decreases
- Volume increases
- Shape becomes indefinite

5. Natural Phenomena (Application)

- Glaciers melt slowly due to high heat of fusion
- Water requires large energy to boil due to strong hydrogen bonding

📊 Diagram:



✦ **Summary:**

- **ΔHf**: Heat required to convert solid ⟶ liquid
- **ΔHv**: Heat required to convert liquid ⟶ gas
- ΔHv > ΔHf because more energy is needed to completely separate molecules
- During phase change ⟶ temperature remains constant
- Energy is used to break intermolecular forces
- Phase changes increase particle motion, distance, and freedom

🌟 **Q5. Differentiate between crystalline and amorphous solids. Explain crystal lattice, properties of solids, and applications of liquid crystals.**

❖ **Answer:**

Introduction

Solids are substances in which particles are closely packed and held by strong forces. Based on the arrangement of particles, solids are classified into crystalline and amorphous solids. Their structure determines their properties and applications.

1. Crystalline Solids

Definition:

Crystalline solids are those in which particles are arranged in a regular, repeating three-dimensional pattern.

Properties:

- **Definite shape:** Due to orderly arrangement
- **Sharp melting point:** Melt at a fixed temperature
- **Long-range order:** Regular structure throughout
- **Anisotropic nature:** Properties vary in different directions
- **Cleavage planes:** Break along specific planes

Examples:

- NaCl, diamond, ice

2. Amorphous Solids

Definition:

Amorphous solids are those in which particles are arranged in an irregular, random manner.

Properties:

- No definite shape
- No sharp melting point (soften gradually)
- Short-range order only
- Isotropic nature: Same properties in all directions
- No cleavage planes

Examples:

- Glass, rubber, plastics

3. Difference Between Crystalline and Amorphous Solids

Arrangement:

- Crystalline $\longrightarrow$ regular
- Amorphous $\longrightarrow$ irregular

Melting point:

- Crystalline $\longrightarrow$ sharp
- Amorphous $\longrightarrow$ gradual

Nature:

- Crystalline ⟶ anisotropic
- Amorphous ⟶ isotropic

Cleavage:

- Crystalline ⟶ present
- Amorphous ⟶ absent

4. Crystal Lattice

Definition:

A crystal lattice is a regular three-dimensional arrangement of atoms, ions, or molecules in a crystalline solid.

Explanation:

- Particles occupy fixed positions
- Pattern repeats throughout the structure
- Determines the shape and properties of the crystal

Importance:

- Gives solids their definite shape
- Explains melting point and strength
- Helps understand physical properties

5. General Properties of Solids

- Definite shape and volume
- High density (particles closely packed)

- Very low compressibility
- Particles vibrate about fixed positions
- Strong intermolecular forces

6. Liquid Crystals

Definition:

Liquid crystals are substances that have properties of both liquids and solids.

Explanation:

- Molecules can flow like liquids
- But remain partially ordered like solids

Properties:

- Flow like liquids
- Show ordered arrangement
- Exhibit optical properties like crystals

7. Applications of Liquid Crystals

1. LCD Screens

- Used in TVs, mobiles, computers
- Control light to form images

2. Medical Diagnostics

- Detect temperature differences

- Used in early detection of tumors

3. Temperature Sensors

- Detect heat changes in circuits

Conclusion

The structure of solids determines their properties. Crystalline solids have ordered structures, while amorphous solids are disordered. Liquid crystals show unique properties and have important modern applications.

✦ **Summary:**

- **Crystalline solids:** Regular structure, sharp melting point
- **Amorphous solids:** Irregular structure, no fixed melting point
- **Crystal lattice:** 3D arrangement of particles
- Solids have strong forces and fixed positions
- **Liquid crystals:** Intermediate state (liquid + solid properties)
- Used in LCDs, medical diagnostics, sensors

📘Exercise

Q.1 Four choices are given for each question. Select the correct choice.

1. London dispersion forces are the only forces present among:

(a) Molecules of water in liquid state

(b) Atoms of helium in gaseous state at high temperature ✅

(c) Molecules of solid iodine

(d) Molecules of hydrogen chloride gas

2. When the vapour pressure of a liquid equals the external pressure, what phenomenon occurs?

(a) Sublimation

(b) Condensation

(c) Boiling ✅

(d) Freezing

3. When water freezes at 0°C its density decreases due to:

(a) Cubic structure of ice

(b) Empty spaces present in the structure of ice ✅

(c) Decrease in volume

(d) Decrease in viscosity

4. Which of the following is the correct sequence of increasing ΔH_v values?

(a) $H_2O > NH_3 > F_2$

(b) $F_2 > NH_3 > H_2O$ ✅

(c) $NH_3 > H_2O > F_2$

(d) $H_2O > F_2 > NH_3$

5. Surface tension of a liquid is due to:

(a) Inward pull of surface molecules ✅

(b) Upward pull from the surface

(c) Collision of molecules

(d) Repulsive forces

6. Which change of state involves overcoming only London dispersion forces?

(a) Melting of ice

(b) Boiling of ethanol

(c) Sublimation of iodine ($I_2(s) \rightarrow I_2(g)$) ✅

(d) Dissolving sodium chloride in water

7. In which substance are London dispersion forces the only intermolecular forces present?

(a) Ammonia (NH_3)

(b) Water (H_2O)

(c) Methane (CH_4) ✅

(d) Hydrogen fluoride (HF)

8. Which is a characteristic property of crystalline solids?

(a) They have a range of melting points

(b) They are isotropic

(c) They have a definite and sharp melting point ✅

(d) They lack regular arrangement

9. Which liquid has the highest viscosity?

(a) Water (H_2O)

(b) Ethanol (CH_3CH_2OH)

(c) Diethyl ether

(d) Glycerol ($CH_2OH–CHOH–CH_2OH$) ✅

10. Which intermolecular force is present in all molecules?

(a) Dipole-dipole forces

(b) Hydrogen bonds

(c) London dispersion forces ✅

(d) Ion-dipole forces

11. Liquid crystals exhibit properties:

(a) Only like solids

(b) Only like liquids

(c) Between solids and liquids ✅

(d) Unlike solids or liquids

Q.2 Attempt the following short-answer short-answer questions

a. Explain, at the molecular level, why evaporation leads to a cooling effect.

Ans: Evaporation causes cooling because high-energy (fast-moving) molecules escape from the surface of a liquid. These molecules carry away more kinetic energy than average. As a result, the remaining liquid has lower average kinetic energy, which means its temperature decreases. This loss of energy from the liquid produces a cooling effect.

b. Explain why liquids with stronger intermolecular forces tend to have lower rates of evaporation at a given temperature compared to liquids with weaker intermolecular forces.

Ans: Liquids with stronger intermolecular forces hold their molecules more tightly together. Therefore, molecules require more energy to escape from the liquid surface. At a given temperature, fewer molecules have enough energy to overcome these strong forces, resulting in a lower rate of evaporation. In contrast, liquids with weaker intermolecular forces allow molecules to escape more easily, leading to faster evaporation.

c. One feels a sense of cooling under the fan after a bath. Explain.

Ans: After a bath, the body has a thin layer of water on its surface. When air from a fan blows over the body, it increases the rate of evaporation of this water. Evaporation requires energy, which is taken from the body in the form of heat. As heat is removed from the body, the body temperature decreases, producing a cooling sensation.

d. How is a dynamic equilibrium established during evaporation of a liquid in a closed vessel at constant temperature?

Ans: In a closed vessel, evaporation and condensation occur simultaneously. Initially, evaporation dominates as molecules escape from the liquid surface. Over time, vapour accumulates above the liquid and some molecules return to the liquid phase (condensation). Eventually, the rate of evaporation becomes equal to the rate of condensation. At this point, a dynamic equilibrium is established, where the number of molecules leaving the liquid equals the number returning, and the vapour pressure remains constant at a given temperature.

e. The boiling point of water is different at Lahore and Murree hills. Explain.

Answer:

Boiling point depends on external atmospheric pressure.

- In Lahore (low altitude), atmospheric pressure is higher ⟶ water boils at a higher temperature.
- In Murree hills (high altitude), atmospheric pressure is lower ⟶ water boils at a lower temperature.

👉 Lower pressure requires less heat for vapour pressure to equal external pressure, so boiling occurs earlier.

f. Discuss two significant consequences of the lower density of ice compared to liquid water in natural environments.

Answer:

1. Ice floats on water:

- Since ice is less dense, it remains on the surface of lakes and rivers.

2. Insulation of aquatic life:

- The floating ice forms a layer on top, preventing the water below from freezing completely. This protects aquatic organisms from extreme cold.

g. Why does the boiling point of a liquid increase when external pressure rises?

Answer:

- Boiling occurs when vapour pressure equals external pressure.

- When external pressure increases, the liquid must reach a higher vapour pressure to boil. This requires more heat, so the boiling point increases.

h. Mention four items in which liquid crystals are used.

Answer:

1. Liquid crystal displays (LCDs) in TVs
2. Mobile phone screens
3. Digital watches and calculators
4. Temperature sensors / thermometers

i. How do you differentiate between crystalline solids and amorphous solids?

Answer:

Arrangement:

- Crystalline ⟶ regular, ordered arrangement
- Amorphous ⟶ irregular, random arrangement

Melting point:

- Crystalline ⟶ sharp melting point
- Amorphous ⟶ no fixed melting point (gradual softening)

Nature:

- Crystalline ⟶ anisotropic (properties differ in directions)
- Amorphous ⟶ isotropic (same properties in all directions)

Cleavage:

- Crystalline ⟶ break along definite planes
- Amorphous ⟶ break irregularly

j. Propanone (CH_3COCH_3), propanol ($CH_3CH_2CH_2OH$), and butane ($CH_3CH_2CH_2CH_3$) have very similar relative molecular masses. Arrange them in increasing order of boiling points and explain your answer.

Answer:

Butane < Propanone < Propanol

Explanation:

Butane is non-polar and has only London dispersion forces, which are weak, so it has the lowest boiling point. Propanone is polar due to the carbonyl group and exhibits dipole-dipole forces, so its boiling point is higher than butane. Propanol contains a hydroxyl group (–OH) and forms hydrogen bonds, which are the strongest intermolecular forces among the three, so it has the highest boiling point.

k. Explain how hydrogen bonding is responsible for the relatively high surface tension of water.

Answer:

Water molecules are strongly attracted to each other through hydrogen bonding. At the surface, molecules experience an inward pull because they are attracted by neighboring molecules beneath them. This creates a tightly packed surface layer, increasing cohesion between molecules at the surface. As a result, water exhibits relatively high surface tension.

l. Identify the dominant intermolecular forces present in the following substances:

Answer:

(i) NH_3 ⟶ Hydrogen bonding

(ii) Ar ⟶ London dispersion forces

(iii) CH_3COCH_3 ⟶ Dipole-dipole forces

(iv) CH_3OH ⟶ Hydrogen bonding

m. The boiling points and molar masses of hydrides of some first-row elements are tabulated below. Suggest reasons for the difference in their boiling points in terms of the type of molecules involved and the nature of the forces present between them.

Table:

Substance	**Boiling Point (K)**	**Molar Mass (g mol^{-1})**
CH_4	112	16
NH_3	240	17
H_2O	373	18

Answer:

The differences in boiling points of CH_4, NH_3, and H_2O are mainly due to the type and strength of intermolecular forces present, rather than their molar masses, which are almost similar.

CH_4 (Methane):

It is a non-polar molecule, so only London dispersion forces act between its molecules. These forces are weak, so CH_4 has the lowest boiling point (112 K).

NH_3 (Ammonia):

It is a polar molecule and forms hydrogen bonds due to N–H bonds and the lone pair on nitrogen. Hydrogen bonding is stronger than dispersion forces, so NH_3 has a higher boiling point (240 K).

H_2O (Water):

It is a highly polar molecule and forms extensive hydrogen bonding networks. Each water molecule can form multiple hydrogen bonds, making intermolecular attraction very strong. Therefore, H_2O has the highest boiling point (373 K).

👉 **Hence**, the boiling points increase in the order:

$CH_4 < NH_3 < H_2O$, due to increasing strength of intermolecular forces.

DESCRIPTIVE QUESTIONS

🌟 **Q.3. What are London dispersion forces? Give examples, and discuss the factors affecting these forces.**

❖ **Answer:**

Introduction

London dispersion forces are the weakest type of intermolecular forces and are present in all atoms and molecules, whether polar or non-polar.

Definition

London dispersion forces are temporary attractive forces that arise due to instantaneous dipoles formed when the electron distribution around atoms or molecules becomes uneven.

Explanation (Molecular Level)

- Electrons in atoms/molecules are constantly moving.
- At any moment, electrons may be unevenly distributed, creating a temporary dipole.
- This temporary dipole can induce a dipole in a nearby atom or molecule.
- The attraction between these temporary dipoles is called London dispersion force.

Examples:

- Noble gases such as He, Ne, Ar

- Non-polar molecules like CH_4 (methane)
- Non-polar substances like I_2 (iodine)

Factors Affecting London Dispersion Forces

1. Size of the Atom or Molecule

- Larger atoms/molecules have more electrons.
- More electrons ⟶ more easily polarizable electron cloud.

👉 **Result**: Stronger dispersion forces

2. Shape of Molecule (Surface Area)

- Linear molecules have more contact area.
- Branched or compact molecules have less contact area.

👉 **Result:** Greater surface area ⟶ stronger dispersion forces

3. Molecular Mass

- Higher molar mass means more electrons.

👉 **Result:** Stronger London dispersion forces

4. Polarizability

- Easily distorted electron clouds increase temporary dipole formation.

👉 **Result:** Higher polarizability ⟶ stronger dispersion forces

Importance

- Responsible for the liquefaction of gases like noble gases
- Affect boiling and melting points of non-polar substances
- Explain physical properties of hydrocarbons and other non-polar compounds

✦ **Summary:**

- London dispersion forces are temporary, weak intermolecular forces present in all substances.
- They arise due to instantaneous dipoles and induced dipoles.

Strength increases with:

- Larger molecular size
- Higher number of electrons
- Greater surface area
- Higher polarizability

They play an important role in determining physical properties like boiling and melting points.

🌟 Q.4 Hydrogen bonding is present in H_2O, NH_3, HF, $(CH_3)_2CO$, and $CHCl_3$ molecules. Explain.

❖ **Answer:**

Introduction

Hydrogen bonding is a special type of intermolecular force that occurs when hydrogen is covalently bonded to a highly electronegative atom such as F, O, or N, and is attracted to a lone pair on another electronegative atom.

✦ Explanation for Each Molecule

1. H_2O (Water)

Water shows strong hydrogen bonding because:

- Oxygen is highly electronegative.
- Each molecule has two lone pairs and can form multiple hydrogen bonds.
- One molecule can act as both hydrogen bond donor and acceptor.

👉 This results in extensive hydrogen bonding networks.

2. NH_3 (Ammonia)

- Nitrogen is electronegative and has a lone pair.
- Hydrogen atoms are bonded to nitrogen (N–H bonds).
- Hydrogen bonding occurs between the hydrogen of one molecule and the lone pair of another.

👉 Hydrogen bonding is present but weaker than in water.

3. HF (Hydrogen Fluoride)

- Fluorine is the most electronegative element.
- Strong polarity of the H–F bond leads to very strong hydrogen bonding.

- Molecules often form chains due to hydrogen bonding.

👉 HF exhibits very strong hydrogen bonding.

4. $(CH_3)_2CO$ (Propanone / Acetone)

- Oxygen in the carbonyl group (C=O) has lone pairs.
- **However**, there is no hydrogen directly bonded to oxygen, nitrogen, or fluorine.

👉 **Therefore:**

- It cannot form hydrogen bonds with itself,
- But can act as a hydrogen bond acceptor in presence of other molecules (like water).

5. $CHCl_3$ (Chloroform)

- Hydrogen is attached to carbon, not to O, N, or F.
- Carbon is not electronegative enough to create strong hydrogen bonding.

👉 **Hence**, $CHCl_3$ does not show significant hydrogen bonding between its own molecules, though weak interactions may occur in mixtures.

Conclusion

Strong hydrogen bonding occurs in H_2O, NH_3, and HF due to highly electronegative atoms (O, N, F).

$(CH_3)_2CO$ and $CHCl_3$ do not exhibit true hydrogen bonding between their own molecules, but acetone can accept hydrogen bonds from other substances.

✦ **Summary:**

- Hydrogen bonding requires H bonded to F, O, or N.
- H_2O, NH_3, HF ⟶ show hydrogen bonding
- $(CH_3)_2CO$ ⟶ only hydrogen bond acceptor
- $CHCl_3$ ⟶ does not show significant hydrogen bonding
- Strength of hydrogen bonding: HF > H_2O > NH_3

🌟 **Q.5 Discuss the structural changes when water turns into ice. Justify the empty spaces in its crystals as compared to H_2O at 4°C and lower density of ice.**

❖ **Answer:**

Introduction

Water shows an unusual property when it freezes. Instead of becoming denser, it becomes less dense due to structural changes in its arrangement of molecules caused by hydrogen bonding.

Structural Changes During Freezing

- In liquid water (especially at 4°C), molecules are closely packed and moving randomly.

- As temperature decreases below 4°C, molecules begin to slow down.
- When water freezes into ice, hydrogen bonds become more stable and fixed, forcing molecules into a rigid crystalline structure.

Structure of Ice

- **In ice,** each water molecule forms four hydrogen bonds with neighboring molecules.
- These bonds arrange molecules in a tetrahedral, hexagonal lattice structure.
- This arrangement is open and spacious, not closely packed.

Reason for Empty Spaces in Ice

- The tetrahedral arrangement creates gaps (empty spaces) between water molecules.
- These empty spaces are due to the specific orientation of hydrogen bonds, which prevent molecules from coming closer together.

As a result, molecules are held in a fixed but less compact structure compared to liquid water.

Comparison with Water at 4°C

- At 4°C, water molecules are packed more closely than in ice.

- Hydrogen bonds are constantly forming and breaking, allowing molecules to occupy less space.

Therefore, water at 4°C has maximum density.

Reason for Lower Density of Ice

- Ice has a larger volume due to empty spaces in its crystal lattice.
- Mass remains the same, but volume increases.
- Since density = mass/volume, an increase in volume results in lower density.

👉 That is why ice is less dense than liquid water and floats on it.

✦ **Summary:**

- Water freezes into a hexagonal lattice structure.
- Hydrogen bonding becomes fixed and forms an open structure.
- Empty spaces appear due to tetrahedral arrangement.
- Water at 4°C is more closely packed than ice.
- Ice has lower density because of increased volume caused by empty spaces.

🌟 Q.6 How liquid crystals resemble liquids and solids? Give their uses in daily life.

❖ **Answer:**

Introduction

Liquid crystals are substances that exist in a state intermediate between solids and liquids. They show properties of both phases due to a partially ordered molecular arrangement.

How Liquid Crystals Resemble Liquids

- They can flow like liquids.
- Their molecules are not fixed in position and can move past each other.
- They do not have a rigid shape and can take the shape of their container.
- They exhibit fluidity similar to liquids.

How Liquid Crystals Resemble Solids

- Their molecules maintain a certain degree of order like solids.
- They show directional arrangement of molecules (orientation is partially fixed).
- They can exhibit optical properties similar to crystals.
- Their structure is more organized than ordinary liquids.

Conclusion

Liquid crystals are intermediate substances that combine the fluidity of liquids with the ordered structure of solids, making them unique in their behavior.

Uses of Liquid Crystals in Daily Life

1. LCD Screens (Liquid Crystal Displays)

- Used in televisions, mobile phones, laptops, and monitors.

2. Digital Watches and Calculators

- Display numbers using liquid crystal technology.

3. Thermometers and Temperature Indicators

- Used to detect temperature changes by color variation.

4. Medical Applications

- Used in medical imaging and temperature-sensitive diagnostics.

5. Optical Devices

- Used in light modulation and display technologies.

✦ Summary:

- Liquid crystals show properties of both liquids (flow) and solids (order).
- They have partial molecular alignment but can still flow.
- Common uses include LCD screens, watches, calculators, thermometers, and medical devices.

🌟 Q.7 Describe the following properties of crystalline solids:

❖ Introduction

Crystalline solids have a highly ordered and repeating arrangement of particles, which gives them definite and characteristic properties.

i) Geometrical Shape

Crystalline solids have a definite and regular geometrical shape because their particles are arranged in a well-ordered three-dimensional pattern. This regular internal structure results in flat faces and specific angles between them.

👉 **Example:** Salt crystals form cubic shapes.

ii) Melting Point

Crystalline solids have a sharp and fixed melting point. This is because all particles are arranged uniformly, and the energy required to break the structure is the same throughout the solid.

👉 They melt suddenly at a specific temperature rather than over a range.

iii) Cleavage Planes

Crystalline solids break along definite planes called cleavage planes. These planes correspond to the natural lines of weakness in the crystal structure where bonds are relatively weaker.

👉 **As a result,** they split into smooth and regular pieces.

iv) Habit of a Crystal

The habit of a crystal refers to its external appearance or characteristic shape under given conditions of growth. It depends on internal atomic arrangement and environmental conditions during crystal formation.

👉 Examples of crystal habits include cubic, prismatic, needle-like, etc.

✦ **Summary:**

- Crystalline solids have a regular geometrical shape due to ordered structure.
- They have a sharp melting point because of uniform bonding.
- They break along cleavage planes due to structural weakness in specific directions.
- Crystal habit refers to the external shape formed during growth conditions.