CHEMISTRY

Unit 5

STATES AND PHASES OF MATTER

MCQs • Short Questions • Long Questions

Complete Study Booklet with Answer Keys

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Section 1 | Multiple Choice Questions

Tick the correct option. Correct answers are highlighted in green with a check mark.

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 deg 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 Delta Hv values of substances mentioned?

✓ **(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 ($H_2O(s) \rightarrow H_2O(l)$)
(b) Boiling of ethanol ($CH_3CH_2OH(l) \rightarrow CH_3CH_2OH(g)$)
✓ **(c) Sublimation of iodine ($I_2(s) \rightarrow I_2(g)$)**
(d) Dissolving sodium chloride in water

7. In which of the following substances are London dispersion forces the only significant intermolecular forces present?

(a) Ammonia (NH_3)
(b) Water (H_2O)
✓ **(c) Methane (CH_4)**
(d) Hydrogen fluoride (HF)

8. Which of the following 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 a regular arrangement of particles

9. Which of the following liquids would you expect to have the highest viscosity at a given temperature?

(a) Water (H2O)
(b) Ethanol (CH3CH2OH)
(c) Diethyl ether (CH3CH2OCH2CH3)
✓ **(d) Glycerol (CH2(OH)CH(OH)CH2(OH))**

10. Which type of intermolecular force is present in all types of molecules regardless of their polarity?

(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

12. Which state of matter exists only within a narrow range of temperature and pressure?

(a) Solid
(b) Gas
✓ **(c) Liquid**
(d) Plasma

13. Why can gases be compressed easily?

(a) Because they have strong intermolecular forces
(b) Because their particles are tightly packed
✓ **(c) Because most of their volume is empty space**
(d) Because they are denser than solids

14. Which one of the following effects explains the cooling during the sudden expansion of a gas?

(a) Boyle's Effect
✓ **(b) Joule-Thomson Effect**
(c) Charles' Effect
(d) Avogadro's Law

15. What is the relation between volume and pressure according to Boyle's Law?

(a) V is proportional to T
(b) V is proportional to n
(c) V is proportional to P
✓ **(d) V is proportional to 1/P**

16. Which of the following statements about ideal gases is true?

(a) They have strong intermolecular forces
(b) Their particles have significant volume
(c) Their volume is mainly due to particle size
✓ **(d) They have negligible intermolecular forces**

17. According to Avogadro's Law, volume is directly proportional to:

(a) Pressure

(b) Temperature
✓ **(c) Number of moles**
(d) Density

18. What is the unit of the ideal gas constant 'R' when pressure is in atm and volume in dm3?
(a) J mol■1 K■1
(b) atm L mol■1 K■1
✓ **(c) atm dm3 mol■1 K■1**
(d) Pa m3 mol■1 K■1

19. The ideal gas equation is given as:
(a) PV = RT
✓ **(b) PV = nRT**
(c) P = nRT/V
(d) n/V = RT/P

20. What is the rearranged form of the ideal gas equation to calculate molar mass (M) of a gas?
(a) M = PV/MRT
✓ **(b) M = mRT/PV**
(c) M = PVnRT
(d) M = RT/PVm

21. Which of the following statements is true about diffusion in liquids?
(a) Liquids do not diffuse at all
(b) Liquids diffuse faster than gases
✓ **(c) Liquids diffuse slower than gases due to stronger intermolecular forces**
(d) Liquids diffuse only when heated

22. What causes the lower compressibility of liquids compared to gases?
(a) Absence of intermolecular forces
(b) Larger molecular size
✓ **(c) Stronger intermolecular forces and less empty space**
(d) High kinetic energy of molecules

23. Liquids are approximately how many times less compressible than gases?
(a) 10
(b) 100
✓ **(c) 105**
(d) 1000

24. According to kinetic molecular theory, liquid molecules are:
(a) At rest
(b) Vibrating in fixed positions
✓ **(c) In constant random motion**
(d) Only moving when heated

25. The expansion of liquids with increase in temperature is:
(a) Equal to that of gases
(b) Greater than that of gases
✓ **(c) Negligible compared to gases**
(d) Same as solids

26. What happens to the kinetic energy of liquid molecules when cooled?
(a) Increases
(b) Remains unchanged

✓ **(c) Decreases, and the liquid may turn into solid**
(d) Becomes zero

27. Which of the following halogens has the highest boiling point?
(a) F2
(b) Cl2
(c) Br2
✓ **(d) I2**

28. Which hydrocarbon has a higher boiling point due to longer chain length?
(a) C2H6
(b) C3H8
✓ **(c) C6H14**
(d) CH4

29. Which of the following has the lowest boiling point among pentane isomers?
(a) n-Pentane
(b) 2-Methylbutane
✓ **(c) 2,2-Dimethylpropane**
(d) Isohexane

30. Which of the following compounds exhibits permanent dipole-permanent dipole interactions?
✓ **(a) HCl**
(b) Cl2
(c) CH4
(d) O2

31. Which of the following is a polar molecule that shows permanent dipole interactions?
(a) CH4
✓ **(b) CHCl3**
(c) Cl2
(d) N2

32. What causes the polarity in a molecule like CHCl3?
(a) Equal sharing of electrons
(b) Non-polar bonds
✓ **(c) Uneven distribution of electrons creating positive and negative centers**
(d) Absence of hydrogen atoms

33. What type of force is hydrogen bonding?
(a) Ionic bond
✓ **(b) A special type of dipole-dipole force**
(c) Metallic bond
(d) London dispersion force

34. Which of the following is essential for hydrogen bond formation?
(a) Hydrogen bonded to a metal
(b) Hydrogen bonded to a non-polar atom
✓ **(c) Hydrogen bonded to a highly electronegative atom (F, O, or N)**
(d) Presence of π-bond

35. What else is required on the electronegative atom for hydrogen bonding besides bonding to H?
(a) Resonance

✓ **(b) Lone pair of electrons**
(c) Double bond
(d) Positive charge

36. Why can a water molecule form two hydrogen bonds on average?
(a) Because it is polar
(b) Due to its temperature
✓ **(c) It has 2 hydrogen atoms and 2 lone pairs**
(d) Because it has a covalent bond

37. Why is HF less acidic compared to HCl, HBr, and HI?
(a) Because F is less electronegative
(b) Because it is non-polar
✓ **(c) Because hydrogen bonding in HF traps the H■ ion**
(d) Because HF is a gas

38. What happens to the volume of water when it freezes?
(a) It contracts by 9%
(b) It remains the same
✓ **(c) It expands by 9%**
(d) It evaporates completely

39. Why does ice float on water?
(a) It is heavier than water
(b) It forms ionic bonds
(c) It is denser than water
✓ **(d) It has lower density than water**

40. What type of structure do water molecules form in ice?
(a) Linear
(b) Hexagonal close-packed
(c) Irregular amorphous
✓ **(d) Regular tetrahedral**

41. What property of water allows aquatic life to survive under frozen surfaces?
(a) High surface tension
(b) High heat capacity
✓ **(c) Ice floats and insulates water below**
(d) Low boiling point

42. What is the value of water's enthalpy of vaporization?
(a) 21 kJ/mol
(b) 30 kJ/mol
✓ **(c) 41 kJ/mol**
(d) 60 kJ/mol

43. Why is water's surface tension so high?
(a) Due to its ionic nature
✓ **(b) Because of hydrogen bonding pulling surface molecules downward**
(c) Because it has low boiling point
(d) Because water molecules are very large

44. Which of the following contributes to water's high viscosity?
(a) Covalent bonds
(b) Metallic interactions

✓ **(c) Strong hydrogen bonding**
(d) Low surface area

45. What causes surface tension in liquids?
(a) Gravity
(b) Air pressure
✓ **(c) Intermolecular forces between surface molecules**
(d) Temperature difference

46. Surface tension usually _______ with increasing temperature.
(a) Increases
✓ **(b) Decreases**
(c) Remains constant
(d) Doubles

47. Why do liquid droplets tend to be spherical in shape?
(a) Due to viscosity
(b) Due to gravity
✓ **(c) Due to surface tension minimizing surface area**
(d) Due to evaporation

48. What is the SI unit of viscosity?
(a) Pascal
(b) Nm■2
✓ **(c) kg m■1 s■1**
(d) Joule

49. Which of the following has greater viscosity than water?
(a) Hexane
✓ **(b) Glycerine**
(c) Acetone
(d) Methanol

50. Earthenware pots keep water cool because:
(a) They are made of clay
(b) They reflect sunlight
(c) They absorb water
✓ **(d) They allow water to evaporate through pores**

51. What is vapour pressure?
(a) Atmospheric pressure
(b) Pressure exerted by a liquid at 0 deg C
✓ **(c) Pressure exerted by vapours in equilibrium with their liquid at a given temperature**
(d) Pressure inside the liquid only

52. At equilibrium in a closed system, which two processes occur at the same rate?
(a) Melting and freezing
(b) Evaporation and boiling
✓ **(c) Evaporation and condensation**
(d) Sublimation and condensation

53. Why does acetone feel colder than water when poured on skin?
(a) It reacts with skin
✓ **(b) It has lower boiling point and evaporates faster**
(c) It is more viscous

(d) It absorbs moisture

54. When does a liquid start boiling?

(a) When temperature reaches 100 deg C
(b) When all molecules turn into vapour
✓ **(c) When its vapour pressure equals atmospheric pressure**
(d) When surface tension becomes zero

55. What is the effect of decreasing external pressure on boiling point of a liquid?

(a) It increases
(b) It remains constant
✓ **(c) It decreases**
(d) It fluctuates

56. Why does water boil at 69 deg C on Mount Everest?

(a) Due to high temperature
(b) Due to low vapour pressure
✓ **(c) Due to low external pressure**
(d) Due to presence of snow

57. What is the role of a pressure cooker in boiling water?

(a) It lowers atmospheric pressure
(b) It allows water to evaporate quickly
✓ **(c) It increases external pressure, raising boiling point**
(d) It cools the steam

58. The value of molar heat of fusion for water is:

(a) 40.6 kJ mol■1
(b) 21.7 kJ mol■1
✓ **(c) 4.6 kJ mol■1**
(d) 15.6 kJ mol■1

59. Which of the following has the lowest molar heat of vaporization?

(a) Water
(b) NH3
✓ **(c) HCl**
(d) Ethanol

60. How does heat of vaporization affect matter?

(a) Causes decrease in kinetic energy
(b) Strengthens intermolecular forces
✓ **(c) Increases particle motion and separates them into gas phase**
(d) Decreases freedom of particles

61. Why is the compression of solids not possible?

(a) Their particles are widely spaced
✓ **(b) The particles are fixed in place and cannot move closer**
(c) The particles are charged
(d) Solids are made of gases.

62. What happens to the particles in a solid when the temperature increases?

(a) They move closer to each other
✓ **(b) Their vibrational motion increases**
(c) They start moving freely
(d) The solid melts

63. In solids, the particles mainly have which type of motion?

(a) Translatory motion
(b) Rotational motion
✓ **(c) Vibrational motion**
(d) Random motion

64. Which of the following is a characteristic of crystalline solids?

(a) They do not have a definite geometrical shape
(b) They melt over a wide temperature range
✓ **(c) They have a sharp melting point**
(d) Their properties do not depend on the direction of measurement

65. When a crystalline solid is broken, it does so along specific planes. These planes are known as:

✓ **(a) Cleavage planes**
(b) Crystal faces
(c) Surface planes
(d) Growth planes

66. The shape in which a crystal usually grows is called its:

✓ **(a) Habit**
(b) Cleavage plane
(c) Crystal lattice
(d) Melting point

67. Amorphous solids are typically:

(a) Rigid and hard
(b) Always crystalline in structure
✓ **(c) Found in lump or fine powder form**
(d) Have a sharp melting point

68. Which of the following materials is an example of an amorphous solid?

(a) Diamond
(b) Ice
✓ **(c) Glass**
(d) Sodium chloride

69. What happens to the shape of a NaCl crystal when 10% urea is present in its solution?

(a) It becomes larger
✓ **(b) It becomes needle-like**
(c) It remains the same
(d) It becomes cubic

70. Which of the following best describes the molecular arrangement in liquid crystals?

(a) Random and disordered
(b) Fixed in all directions
✓ **(c) Parallel and partially ordered**
(d) Completely rigid and fixed

71. What type of molecular shape is commonly found in liquid crystals?

(a) Spherical
(b) Cubic
✓ **(c) Rod-like and elongated**
(d) Spiral

72. Liquid crystals exhibit properties of:

(a) Only solids
(b) Only liquids
✓ **(c) Both solids and liquids**
(d) Gases and solids

73. Which of the following is a unique optical property of liquid crystals?

(a) They reflect ultraviolet light
(b) They are always opaque
(c) They are isotropic
✓ **(d) They are anisotropic**

74. In which device are liquid crystals most commonly used?

(a) Generators
✓ **(b) LCD screens**
(c) Loudspeakers
(d) Solar panels

75. Why are cholesteric liquid crystals useful in medical diagnostics?

(a) They kill bacteria
(b) They react with tumors chemically
✓ **(c) They detect temperature differences**
(d) They absorb X-rays

76. Liquid crystals are used in thermometers because:

(a) They are highly reflective
(b) They glow at high temperatures
✓ **(c) Their color changes with temperature**
(d) They conduct electricity

77. Which of the following is NOT a property of liquid crystals?

(a) Fluidity
(b) Rigidity like crystals
✓ **(c) Complete isotropy**
(d) Viscosity

Section 2 | Short Questions

Q1. Explain, at the molecular level, why evaporation leads to a cooling effect.

In a liquid, molecules move with varying kinetic energies - some fast, some slow.

Molecules with higher-than-average kinetic energy overcome intermolecular forces and escape the surface as vapour, taking extra energy with them.

This lowers the average kinetic energy of the remaining molecules, so the liquid's temperature drops, producing a cooling effect.

Q2. Why do liquids with stronger intermolecular forces have lower rates of evaporation than liquids with weaker forces, at a given temperature?

A molecule needs enough kinetic energy to overcome intermolecular attraction before it can evaporate.

In liquids like water (strong hydrogen bonds) only a few molecules have enough energy to escape at any moment, so evaporation is slow.

In liquids like acetone (weaker dipole-dipole/dispersion forces) more molecules can escape at the same temperature, so evaporation is faster.

Q3. One feels a sense of cooling under a fan after taking a bath. Explain.

Water on the skin evaporates, and evaporation requires energy which is taken from body heat.

The most energetic water molecules escape first, lowering the average kinetic energy (and temperature) of the skin.

The fan blows away humid air and replaces it with drier air, speeding up evaporation and heat removal, which feels cooling.

Q4. How is dynamic equilibrium established during evaporation of a liquid in a closed vessel at constant temperature?

Molecules evaporate from the liquid; some vapour molecules collide with the surface and condense back.

Over time, the rate of evaporation becomes equal to the rate of condensation - this is dynamic equilibrium.

At equilibrium, the amount of liquid and vapour remains constant and the vapour pressure becomes stable.

Q5. The boiling point of water is different at Lahore and Murree hills. Why?

Boiling point depends on atmospheric pressure.

Lahore is at lower altitude with higher atmospheric pressure, so water boils at 100 deg C; Murree Hills are at higher altitude with lower pressure, so water boils at a lower temperature.

Q6. Discuss two significant consequences of the lower density of ice compared to liquid water.

Floating of Ice: Ice floats and forms an insulating layer on lakes/rivers in winter, preventing the water below from freezing and helping aquatic life survive.

Weathering of Rocks: Water entering rock cracks expands on freezing (lower density), exerting pressure that breaks rocks apart over time - frost weathering.

Q7. Why does the boiling point of a liquid increase when external pressure rises?

Boiling occurs when vapour pressure equals external pressure.

When external pressure increases, the liquid must reach a higher temperature for its vapour pressure to match it, so the boiling point rises.

Q8. Mention four items in which liquid crystals are used.

Liquid crystal displays (LCDs) in TV screens; computer monitors and laptop displays; cell phone displays; calculators and watches.

Additional uses: diagnostic devices to detect tumors/infections, and temperature sensors/thermometers.

Q9. How do you differentiate between crystalline solids and amorphous solids?

Crystalline solids have a definite geometric shape, sharp melting point, and long-range ordered arrangement (anisotropic).

Amorphous solids lack a regular pattern, melt over a range of temperatures, and are isotropic (glass, plastic, etc.).

Q10. Propanone, propanol and butane have similar molecular masses. Arrange them in increasing order of boiling point and explain.

Order: Butane < Propanone < Propanol.

Butane has only weak London dispersion forces - lowest boiling point.

Propanone has dipole-dipole interactions from its polar carbonyl group - stronger than London forces.

Propanol has hydrogen bonding via its -OH group - the strongest force, giving the highest boiling point.

Q11. Discuss how hydrogen bonding is responsible for the relatively high surface tension of water.

Water molecules form strong hydrogen bonds due to attraction between partially positive H and partially negative O atoms.

Surface molecules experience unbalanced inward forces, pulling them together and creating a tightly held surface layer - this produces high surface tension.

Q13. Boiling points of hydrides of first-row elements differ due to molecule type and intermolecular forces. Explain for CH_4, NH_3, H_2O.

CH_4 (b.p. 112 K): Nonpolar, only weak London dispersion forces - lowest boiling point.

NH_3 (b.p. 240 K): Polar, forms hydrogen bonds via N-H bonds - stronger than dispersion forces.

H_2O (b.p. 373 K): Extensive hydrogen bonding via two O-H bonds and high oxygen electronegativity - highest boiling point.

Order of strength: London Forces < Dipole-Dipole < Hydrogen Bonding.

Q14. Why is the rate of diffusion in liquids lower than in gases?

Intermolecular forces in liquids are stronger, restricting free movement of molecules.

Liquid molecules are closely packed, leaving little space for fast diffusion compared to gases.

Q15. What happens to the volume of a gas when pressure is doubled?

At constant temperature, doubling the pressure halves the volume, per Boyle's Law (V is proportional to 1/P).

Gases are highly compressible due to the large empty space between molecules.

Q16. How do liquid molecules behave according to kinetic molecular theory?

Liquid molecules are in constant random motion with kinetic energy that increases with temperature.

This motion allows diffusion and flow, though slower than in gases due to stronger intermolecular forces.

Q17. Why are liquids able to flow?

Liquid molecules are not fixed in place and can move past each other.

Although closely packed, the intermolecular forces are not strong enough to lock them completely, allowing fluid motion.

Q18. How does increasing temperature affect the volume of a liquid?

Increased temperature raises the kinetic energy of liquid molecules, causing slight expansion.

This expansion is much smaller than in gases since liquid molecules are already close together with stronger intermolecular forces.

Q19. Why do non-polar molecules exhibit intermolecular forces despite lacking permanent dipoles?

Non-polar molecules exhibit instantaneous dipole-induced dipole forces due to temporary shifts in electron clouds, creating momentary dipoles that induce dipoles in neighbouring molecules.

Q20. How does increasing molecular mass affect the boiling point of noble gases and halogens?

With increasing molecular mass, size and polarizability of atoms increase, strengthening id-id (dispersion) forces and raising boiling points down a group.

Q21. Why is iodine a solid and fluorine a gas at room temperature?

Iodine has much higher molecular mass and size than fluorine, giving stronger id-id forces and making it a solid, while fluorine's weaker forces keep it a gas.

Q22. What are permanent dipole-permanent dipole forces?

Forces of attraction between the positive end of one polar molecule and the negative end of another, occurring in molecules like HCl and CHCl3 that carry permanent partial charges.

Q23. How does the shape or surface area of a hydrocarbon molecule affect its boiling point?

Molecules with larger or linear surface area (e.g. n-pentane) have stronger id-id forces and higher boiling points.

Branched molecules have smaller surface areas and weaker forces, hence lower boiling points.

Q24. Why does hexane have a higher boiling point than ethane?

Hexane has greater molecular mass and a longer chain, giving stronger instantaneous dipole-induced dipole forces than ethane, hence a higher boiling point.

Q25. What is boiling point and when does a liquid boil?

Boiling point is the temperature at which a liquid's vapour pressure equals atmospheric pressure.

At this point bubbles form and rise continuously from within the liquid.

Q26. Why is the boiling point of water higher than alcohol?

Water has strong hydrogen bonding, requiring more energy to overcome these forces, giving a higher boiling point than alcohols.

Q27. Why does water boil at a lower temperature on Mount Everest?

Atmospheric pressure is lower on Mount Everest, so water reaches its vapour pressure at a lower temperature and boils at only 69 deg C.

Q28. What conditions are necessary for hydrogen bonding to occur?

A hydrogen atom must be covalently bonded to a highly electronegative atom (F, O, or N).

That electronegative atom must have a lone pair of electrons to attract the partially positive hydrogen.

Q29. Why is hydrogen bonding considered stronger than other intermolecular forces?

Hydrogen bonding involves highly polarized covalent bonds; the strong attraction between delta+ hydrogen and delta- electronegative atoms produces a significant intermolecular force.

Q30. Why does water form more hydrogen bonds than ammonia or HF?

Water has two hydrogen atoms and two lone pairs on oxygen, enabling two hydrogen bonds per molecule.

NH3 has only one lone pair and HF has only one H atom, limiting their bonding capacity.

Q31. How does hydrogen bonding affect the boiling point of water?

Extensive hydrogen bonding among water molecules requires a large amount of energy to break, making water's boiling point (100 deg C) unusually high compared to other Group 16 hydrides.

Q32. Why is ice less dense than liquid water?

In ice, hydrogen bonding arranges water molecules in a tetrahedral structure with empty spaces, increasing volume and decreasing density, so ice floats.

Q33. What is the role of hydrogen bonding in the high heat capacity of water?

Hydrogen bonding lets water absorb and store more heat energy without a large temperature rise, giving it a high specific heat capacity that stabilizes temperature in environments.

Q34. How does hydrogen bonding explain the high surface tension of water?

Surface water molecules are strongly attracted to those beneath via hydrogen bonding, creating a stretched surface film responsible for high surface tension.

Q35. What is viscosity and how is it affected by hydrogen bonding in water?

Viscosity is a liquid's resistance to flow. Water's strong hydrogen bonds create internal friction, increasing its viscosity compared to non-hydrogen-bonding liquids like hexane.

Q36. Why does HF have a higher boiling point than NH3 despite forming only one bond?

HF forms strong hydrogen bonds due to fluorine's high electronegativity; although it forms only one bond per molecule, its strength raises the boiling point above ammonia's.

Q37. What is surface tension and why does it occur in liquids?

Surface tension is the force acting along a liquid's surface, making it behave like a stretched elastic sheet.

It arises from unbalanced inward forces on surface molecules caused by intermolecular attraction.

Q38. Why do water droplets form spherical shapes?

Surface tension tries to minimize surface area; a sphere has the smallest surface area for a given volume, making it the most energy-efficient shape.

Q39. How does temperature affect surface tension?

As temperature increases, molecular kinetic energy increases, reducing intermolecular attraction, so surface tension decreases with rising temperature.

Q40. Why does water have higher surface tension than alcohol or acetone?

Water has strong hydrogen bonds giving high intermolecular forces; alcohol and acetone have weaker forces, hence lower surface tension.

Q41. What is viscosity in liquids?

Viscosity is a liquid's resistance to flow, arising from friction between adjacent layers where slower outer layers resist the movement of faster inner ones.

Q42. How does temperature affect viscosity of a liquid?

With increasing temperature, molecular motion becomes faster and weakens intermolecular attractions, lowering viscosity and making the liquid flow more easily.

Q43. Why is honey more viscous than water?

Honey has stronger intermolecular forces and a more complex molecular structure than water, giving greater resistance to flow and higher viscosity.

Q44. Why is glycerine more viscous than hexane?

Glycerine forms hydrogen bonds via its hydroxyl groups, while hexane is non-polar with no such interactions; the stronger forces in glycerine give it higher viscosity.

Q45. What is evaporation and how does it cause cooling?

Evaporation is the escape of high-energy molecules from a liquid's surface, leaving behind low-energy molecules and lowering the liquid's temperature.

Q46. Why do earthenware pots keep water cool?

Earthen pots are porous, letting water slowly evaporate through them; the energy for evaporation comes from the water itself, lowering its temperature.

Q47. What is vapour pressure of a liquid?

Vapour pressure is the pressure exerted by vapour molecules in equilibrium with their liquid at a given temperature, in a closed container where evaporation and condensation rates are equal.

Q48. Why does vapour pressure not depend on the amount of liquid?

Vapour pressure depends only on temperature and intermolecular forces; once equilibrium is reached, the quantity or surface area of liquid does not affect it.

Q49. What happens to the rate of condensation as surface area increases?

A larger surface area means more vapour molecules collide with the liquid surface, increasing the rate of condensation to match the evaporation rate more quickly.

Q50. Why does petrol have higher vapour pressure than kerosene?

Petrol has weaker intermolecular forces and more volatile components than kerosene, giving it a higher vapour pressure.

Q51. How does a pressure cooker increase the boiling point of water?

In a pressure cooker, trapped vapour raises internal pressure, which increases the boiling point of water and allows food to cook faster.

Q52. Why does acetone feel colder on the skin than water?

Acetone evaporates faster due to weaker intermolecular forces; its rapid evaporation absorbs more heat from the skin, giving a greater cooling effect.

Q53. How does temperature affect vapour pressure?

As temperature increases, molecular kinetic energy rises, increasing evaporation and vapour pressure until equilibrium is reached at a higher level.

Q54. What is molar heat of fusion (Delta Hf)?

The amount of heat absorbed by one mole of a solid to convert it into liquid at its melting point under 1 atm pressure.

For water: Delta Hf = 4.6 kJ/mol.

Q55. What is molar heat of vaporization (Delta Hv)?

The amount of heat required to convert one mole of a liquid into vapour at its boiling point under 1 atm pressure.

For water: Delta Hv = 40.6 kJ/mol.

Q56. Why is Delta Hv of water higher than that of NH3 or HCl?

Water has stronger hydrogen bonding, which requires more energy to overcome, compared to the weaker forces in NH3 (21.7 kJ/mol) and HCl (15.6 kJ/mol).

Q57. What happens to particles during fusion and vaporization?

Particles gain energy and move farther apart, changing phase - from solid to liquid during fusion, and from liquid to gas during vaporization.

Q58. Why are solids rigid and incompressible?

Their particles are tightly packed by strong interparticle forces, leaving no significant space for compression.

Q59. What kind of motion do particles in a solid show?

They only vibrate about their mean positions; no translational or rotational motion occurs.

Q60. How does temperature affect solid expansion?

Temperature slightly increases vibrational motion, but expansion is negligible because strong forces keep particles fixed.

Q61. What is the main reason solids have high melting points?

Strong intermolecular or interatomic forces (ionic, covalent, metallic, etc.) require high energy to break, giving solids high melting points.

Q62. What is the kinetic energy of solid particles according to KMT?

Only vibrational kinetic energy, since the particles cannot translate or rotate.

Q63. Why can solids not be compressed like gases?

Solid particles are already closely packed with minimal available space (maximum about 74% space occupied), leaving little room for compression.

Q64. What is the main difference between crystalline and amorphous solids?

Crystalline solids have a regular, definite geometric shape and long-range order of particles.

Amorphous solids lack such regularity and typically appear as lumps or powders.

Q65. What is a crystal lattice?

A crystal lattice is the regular, three-dimensional arrangement of ions, atoms, or molecules in space that forms the structure of crystalline solids.

Q66. What are the properties of crystalline solids?

Definite geometric shape; sharp melting points; cleavage planes; growth through crystallization from a saturated solution; a characteristic habit (shape) during growth.

Q67. What happens when a crystalline solid is broken?

Crystalline solids break along specific planes called cleavage planes, which have a characteristic angle relative to each other.

Q68. How are crystals grown?

Crystals grow by slow evaporation of a solvent or by seeding from a saturated solution, where particles arrange in a specific, orderly fashion.

Q69. What is the habit of a crystal?

The habit of a crystal is the typical shape it takes when it grows under specific conditions; the shape can change if conditions like the presence of impurities change.

Q70. How do amorphous solids differ in structure from crystalline solids?

Amorphous solids do not have a regular, repeating pattern of particles and lack the long-range order characteristic of crystalline solids.

Q71. Do amorphous solids have sharp melting points?

No, amorphous solids melt over a wide range of temperatures rather than at a specific melting point.

Q72. What is the main difference between crystalline and amorphous solids regarding direction?

Crystalline solids are anisotropic (properties depend on the direction of measurement).

Amorphous solids are isotropic (properties are the same in all directions).

Q73. Can amorphous solids be molded into different shapes?

Yes, amorphous solids like glass can be molded and blown into various shapes due to their lack of a definite structure.

Q74. What are liquid crystals?

Liquid crystals are substances that are neither fully liquid nor fully solid; their molecules can move like viscous liquids but have a restricted range of motion like solids.

Q75. What is the shape of the molecules in liquid crystals?

Molecules in liquid crystals are typically rigid, rod-like, and linear, with their length being four to eight times greater than their diameter.

Q76. How do liquid crystals move?

Liquid crystals flow like liquids, but their motion is limited - they can spin around their long axis but cannot rotate end over end.

Q77. How are the molecules in liquid crystals arranged?

Molecules tend to align parallel to each other along their long axis, similar to logs stacked in a pile of firewood.

Q78. What properties of liquid crystals resemble liquids?

Liquid crystals flow like liquids and show viscosity similar to liquids, as they can move and flow to some extent.

Q79. What is the primary application of liquid crystals?

The main application of liquid crystals is in electro-optic devices such as Liquid Crystal Displays (LCDs), which modulate light electrically.

Q80. How is the viscosity of liquid crystals?

The viscosity of liquid crystals is similar to liquids; they show fluid-like behaviour, although their motion is more restricted than typical liquids.

Q81. What optical properties do liquid crystals exhibit?

Liquid crystals show optical properties similar to crystals, reflecting or transmitting light in a specific manner, making them suitable for display technologies.

Q82. How are liquid crystals used to detect tumors?

Tumors often have a different temperature than surrounding tissue; cholesteric liquid crystals applied to the skin can detect this temperature difference to locate tumors.

Q83. Where are liquid crystal temperature sensors used?

They are used to identify faulty connections on circuit boards and in thermometers to detect temperature changes.

Q84. What are some daily applications of liquid crystals?

Liquid crystals are commonly used in electronic displays such as TVs, computer screens, calculators, and watches.

Q85. How are liquid crystals applied in industrial and scientific fields?

Liquid crystals are increasingly important in electro-optic devices used in industry and science, for controlling light and displaying information.

Q86. How is the relative molecular mass (Mr) of a gas calculated?

The general gas equation $PV = nRT$ can be used when P, T, V and mass in grams are known.

Putting $n = m/M$ gives $PV = (m/M)RT$, which rearranges to $M = mRT/PV$; at the molecular level this becomes $M_r = mRT/PV$.

Q87. Drugs with higher viscosity are not oxidized easily - justify.

High-viscosity drugs have reduced diffusion and dissolution due to a complex internal structure, which limits their exposure to oxidizing conditions, so they are not easily oxidized.

Section 3 | Descriptive & Numerical Problems

Descriptive Questions

Q3 (Descriptive). What are London dispersion forces? Give examples and discuss the factors affecting them.

London dispersion forces (instantaneous dipole-induced dipole forces) arise from momentary, temporary dipoles created by shifting electron clouds, which induce dipoles in neighbouring molecules.

They are present in all molecules but are the only forces in non-polar species like F_2, Cl_2, He, Ne.

Factors: (1) Molecular mass and size - larger atoms/molecules have more easily polarizable electron clouds, giving stronger forces (e.g. F_2, Cl_2 gases; Br_2 liquid; I_2 solid). (2) Surface area/shape - molecules with larger surface area (e.g. straight-chain hydrocarbons, n-pentane) have stronger forces and higher boiling points than branched isomers with smaller surface area.

Q4 (Descriptive). Hydrogen bonding is present in H_2O, NH_3, $(CH_3)_2CO$ and $CHCl_3$ molecules. Discuss briefly.

Hydrogen bonding occurs when H is covalently bonded to a highly electronegative atom (F, O, N) that also carries a lone pair.

H_2O: two H atoms and two lone pairs on O allow each molecule to form two hydrogen bonds on average, giving extensive 3D hydrogen bonding and a high boiling point.

NH_3: only one lone pair on N limits it to one hydrogen bond per molecule, giving weaker bonding and a lower boiling point than water.

$(CH_3)_2CO$ (acetone) and $CHCl_3$ (chloroform) are polar but lack an H atom directly bonded to F, O, or N in a position for classic hydrogen bonding; they instead show dipole-dipole interactions (though acetone's carbonyl oxygen can accept hydrogen bonds from suitable donors).

Q5 (Descriptive). Discuss the structural changes when water turns into ice, and explain the empty spaces and lower density of ice.

Water molecules have a tetrahedral arrangement with two lone pairs occupying two corners of the tetrahedron.

When water freezes, the molecules arrange into a more regular hydrogen-bonded lattice that contains empty spaces.

Ice occupies about 9% more volume than liquid water at 4 deg C, so its density decreases and ice floats - an anomalous behaviour that insulates water bodies and helps aquatic life survive winter.

Q6 (Descriptive). How do liquid crystals resemble liquids and solids? Give their uses in daily life.

Like solids, liquid crystal molecules are elongated, rod-like, show parallel ordered arrangement, are somewhat rigid, and are anisotropic.

Like liquids, they flow, show viscosity, and can fill containers.

Uses: LCD screens in TVs, computers, calculators, watches, and phones; medical diagnostics for detecting tumors via temperature differences; temperature sensors and thermometers.

Q7 (Descriptive). Describe: (i) Geometrical shape (ii) Melting point (iii) Cleavage plane (iv) Habit of a crystal, for crystalline solids.

(i) Geometrical shape: crystalline solids have a definite, distinctive geometric shape from the orderly 3D arrangement of particles; interfacial angles stay the same regardless of crystal size.

(ii) Melting point: crystalline solids melt sharply at a definite, characteristic temperature.

(iii) Cleavage plane: when broken, crystalline solids split along definite planes inclined at a characteristic angle.

(iv) Habit: the typical shape a crystal grows into under given conditions; it can change if conditions (e.g. impurities) change - e.g. NaCl grows cubic but becomes needle-like with 10% urea present.

Numerical Problems

Q8. A sample of an unknown gas has a mass of 0.560 g. It occupies 2.87x10^-4 m3 at 300 K and 1.01x10^5 Pa. Calculate the molar mass. (R = 8.314 J/mol/K)

Given: m = 0.560 g; V = 2.87x10^-4 m3; T = 300 K; P = 1.01x10^5 Pa; R = 8.314 Nm/mol/K

M = mRT/PV = (0.560)(8.314)(300) / [(1.01x10^5)(2.87x10^-4)]

M = 1396.8 / 28.99 = 48.18 g/mol

Result: The molar mass of the unknown gas is 48.18 g/mol.

Q9. 150 cm3 of a volatile liquid vaporizes completely at 98 deg C and 1.01x10^5 Pa; vapour mass = 0.495 g. Find the molecular mass. (R = 8.314 J/mol/K)

Given: V = 150 cm3 = 1.5x10^-4 m3; T = 98 deg C = 371 K; P = 1.01x10^5 Pa; m = 0.495 g

M = mRT/PV = (0.495)(8.314)(371) / [(1.01x10^5)(1.5x10^-4)]

M = 1526.82 / 15.15 = 100.78 g/mol

Result: The molar mass of the vapour is 100.78 g/mol.

Section 4 | Long Questions

Q1. How many states of matter exist? Explain the properties of gases and liquids.

States of Matter

Matter exists in four states: Solid, Liquid, Gas, and Plasma.

The gaseous state is the simplest form of matter. Liquids are less common than solids, gases and plasmas because the liquid state of any substance exists only within a relatively narrow range of temperature and pressure.

Properties of Gases

Gases have negligible intermolecular forces and about 99% empty space between particles, making them highly compressible and able to expand to fill any container.

Gas particles move randomly at high speed (translational motion) and diffuse rapidly.

Gases show large expansion with temperature and their volume, pressure and temperature are related by the gas laws.

Properties of Liquids

Liquid particles are closely packed with stronger intermolecular forces than gases, giving liquids a definite volume but no fixed shape (they take the shape of their container).

Liquids are far less compressible than gases (about 10^5 times less) because there is little empty space between particles.

Liquids diffuse and flow, but more slowly than gases due to stronger intermolecular forces; they also expand with temperature, but the expansion is small compared to gases.

Q2. Derive an expression for the ideal gas equation.

Gas Laws

Boyle's Law: V is proportional to 1/P (n, T constant)

Charles's Law: V is proportional to T (n, P constant)

Avogadro's Law: V is proportional to n (P, T constant)

Derivation of the Ideal Gas Equation

Combining the three gas laws: V is proportional to nT/P

Introducing a proportionality constant R (the ideal gas constant): V = R x nT/P

Rearranging gives the ideal gas equation: PV = nRT

Calculation of Relative Molecular Mass (Mr) of a Gas

From PV = nRT, substituting n = m/M gives PV = (m/M)RT, which rearranges to M = mRT/PV.

Sample Problem 5.1

134 g of a gas at -73 deg C (200 K) under 10 atm pressure occupies 5 dm3. Find the relative molecular mass.

M = mRT/PV = (134)(0.0821)(200) / [(10)(5)] = 44 g/mol

Answer: The relative molecular mass of the gas is 44 amu.

Quick Check 5.1

(a) Gases can be compressed easily because about 99% of their volume is empty space between molecules.

(b) Joule-Thomson Effect: when a compressed gas is allowed to expand suddenly, it produces a cooling effect.

(c) 21 g of a gas occupies 8 dm3 at -90 deg C (183 K) under 7 atm. M = mRT/PV = (21)(0.0821)(183) / [(7)(8)] = 5.63 g/mol - the relative molecular mass is 5.63 amu.

Q3. What are intermolecular forces? Discuss permanent dipole-permanent dipole forces (pd-pd) with examples.

Intermolecular Forces

Definition: the forces of attraction among molecules are called intermolecular forces.

Types: (1) Instantaneous dipole-induced dipole forces (London dispersion forces), (2) Permanent dipole-permanent dipole forces, (3) Hydrogen bonding.

Permanent Dipole-Permanent Dipole Forces (pd-pd)

Definition: the force of attraction between the positive end of a polar molecule and the negative end of a nearby polar molecule.

Example - HCl: in the liquid state, HCl molecules line up with the positive end of one near the negative end of another.

Example - CHCl3 (chloroform): has a positive centre at the H atom and a negative centre at the Cl-rich end.

Q4. What are instantaneous dipole-induced dipole forces (id-id)? Explain the factors affecting them.

Instantaneous Dipole-Induced Dipole Forces (id-id)

Definition: the momentary force of attraction between an instantaneous dipole and an induced dipole in a neighbouring molecule.

These forces exist in all molecules but are the only significant forces in non-polar molecules such as F2, Cl2 (halogens) and He, Ne (noble gases).

Factor 1: Molecular Mass and Size

Greater molecular mass (Mr) gives stronger id-id forces because larger atoms/molecules have electron clouds that are more easily polarized/dispersed.

This explains the trend in halogens and noble gases: F2 and Cl2 are gases, Br2 is liquid, and I2 is solid, as size and polarizability increase down the group.

Factor 2: Surface Area (Shape of Molecule)

A molecule with larger surface area has more points of contact with neighbouring molecules, giving stronger forces.

Example: C_2H_6 (ethane) b.p. = -88.6 deg C vs C_6H_{14} (hexane) b.p. = 68.7 deg C - longer molecules have more contact area.

Isomers of pentane (same molecular mass, different surface area): n-Pentane (largest surface area, highest b.p.) > 2-Methylbutane (intermediate) > 2,2-Dimethylpropane (smallest surface area, lowest b.p.).

Quick Check 5.2

(a) Straight-chain hydrocarbons in gasoline are converted to branched-chain ones to improve quality, because straight chains have stronger intermolecular forces (larger surface area) and need more energy to break, whereas branched chains break more easily.

(b) CCl_4 and SiF_4 are non-polar, so only instantaneous dipole-induced dipole (London dispersion) forces act between their molecules.

Q5. What is hydrogen bonding? Explain it in various compounds.

Hydrogen Bonding

Definition: the force of attraction between a partially positively charged hydrogen atom and a highly electronegative atom (F, O, N) on a neighbouring molecule.

Conditions: (1) H is covalently bonded to a highly electronegative atom (F, O, or N). (2) That electronegative atom must have a lone pair of electrons.

It is a special type of dipole-dipole force, the strongest of the intermolecular forces, but weaker than ionic, metallic and covalent bonds (about one-tenth the strength of an ordinary covalent bond). It is represented by a dotted line.

(i) Liquid Water (H_2O)

Water has two H atoms and two lone pairs on oxygen, forming two hydrogen bonds on average per molecule and extensive 3D hydrogen bonding, responsible for its high boiling point (100 deg C).

(ii) Ammonia (NH_3)

NH_3 forms only one hydrogen bond per molecule despite having three H atoms, because nitrogen has only one lone pair - giving weaker hydrogen bonding and a lower boiling point (-33 deg C).

(iii) Hydrogen Fluoride (HF)

F has three lone pairs (potentially allowing three H-bonds), but HF has only one H atom, restricting it to one bond per molecule. Its boiling point (19.9 deg C) is much higher than NH_3's, and it forms zig-zag chains in solid form.

(iv) Hydrogen Halides

The exceptionally low acidic strength of HF compared to HCl, HBr and HI is attributed to strong hydrogen bonding, which traps the partially positive hydrogen between two highly electronegative atoms.

Q6. Explain the effects of hydrogen bonding on the properties of water.

(i) Structure and Low Density of Ice

Water molecules have a tetrahedral structure with two lone pairs occupying two corners.

On freezing, molecules arrange more regularly, creating empty spaces in the hydrogen-bonded lattice.

Ice occupies about 9% more space than liquid water, so its density decreases and it floats - an anomalous behaviour that lets fish survive under frozen lakes, insulated by the ice layer.

(ii) High Heat Capacity

Water has a high specific heat capacity because strong hydrogen bonds absorb and store considerable heat energy.

(iii) Anomalous Heat of Vaporization and Boiling Point

General trend for Group 16 hydrides: $H_2S < H_2Se < H_2Te < H_2Po$ (increasing).

Water is the exception, with the highest enthalpy of vaporization (41 kJ/mol) and a remarkably higher boiling point than the rest of the group's hydrides.

(iv) Surface Tension and Viscosity

Surface Tension: high, due to the downward pull of surface water molecules through hydrogen bonds.

Viscosity: high, due to strong hydrogen bonding creating internal friction between molecules.

Q7. Explain the surface tension and viscosity of liquids.

Surface Tension of Liquids

Definition: the force acting along the surface of a liquid, making it behave like a stretched elastic sheet.

It arises because surface molecules experience a net inward force from intermolecular attraction, so liquids minimize surface area (forming spherical droplets). It influences capillary action, wetting, and bubble/droplet formation.

Factors: surface tension decreases with increasing temperature and increases with stronger intermolecular forces.

Viscosity of Liquids

Definition: viscosity is a liquid's resistance to flow - describing its thickness or stickiness.

Higher viscosity means thicker liquid with more resistance to flow (e.g. honey is more viscous than water).

Units: SI - kg/m/s; CGS - Poise (P).

Factors: viscosity decreases with increasing temperature (molecules overcome attractive forces more easily) and increases with stronger intermolecular forces.

Quick Check 5.4

(a) Increasing order of surface tension: $CH_3OCH_3 < CH_3COCH_3 < C_2H_5OH$ - ethanol has strong hydrogen bonding and low volatility, giving the highest surface tension; acetone is intermediate; dimethyl ether is lowest.

(b) CCl_4 has higher viscosity than $CHCl_3$ (larger molecular size despite weaker London forces vs. $CHCl_3$'s dipole-dipole forces) but lower viscosity than ethanol (which has strong hydrogen bonding).

(c) Honey's viscosity is far higher than water's due to strong intermolecular forces influenced by water content, temperature and composition.

(d) Glycerine is more viscous than hexane because glycerine has hydrogen bonding plus London forces, while hexane has only London forces.

Q8. Write a note on evaporation and vapour pressure.

Evaporation

Definition: the spontaneous conversion of a liquid into vapour at any temperature.

Evaporation causes cooling: high-energy molecules leave the liquid, low-energy molecules remain, the liquid's temperature falls, and heat moves in from the surroundings.

Examples: cooling felt after a bath (water evaporating from the body); earthenware vessels keep water cool as it evaporates through the pores, taking energy from the water.

Vapour Pressure

Definition: the pressure exerted by a liquid's vapour in equilibrium with its liquid at a given temperature.

In a closed system, molecules leaving the surface collide with the container walls and the liquid surface; some are recaptured (condensation). At equilibrium, the rate of evaporation equals the rate of condensation.

Vapour pressure does not depend on the amount of liquid, its volume, or its surface area; it decreases with stronger intermolecular forces and increases with rising temperature.

Quick Check 5.5

(a)(i) Alcohol has higher vapour pressure than glycerine. (ii) Petrol has higher vapour pressure than kerosene. (iii) Water has higher vapour pressure than mercury.

(b) Glycerine is more viscous than kerosene due to hydrogen bonding, while kerosene is non-polar with weak London forces.

(c) Acetone feels colder than water on skin because, being non-polar with weak intermolecular forces, it evaporates faster, absorbing heat quickly.

(d) Evaporation is faster at higher temperature because increased kinetic energy raises vapour pressure and speeds up the process.

(e) We feel cool near a riverbank mainly due to evaporative cooling - evaporating water absorbs heat from the surrounding air, lowering its temperature.

Q9. What is boiling point? Discuss the factors affecting boiling point.

Boiling Point

Definition: the temperature at which a liquid's vapour pressure becomes equal to the external (atmospheric) pressure.

When a liquid is heated, its vapour pressure increases; at the boiling point, vapour pressure equals atmospheric pressure, bubbles form within the liquid and burst at the surface. Further heating does not raise the temperature further - it goes into breaking intermolecular forces.

Factor 1: Strength of Intermolecular Forces

Stronger intermolecular forces give lower vapour pressure at a given temperature and hence a higher boiling point - water's boiling point is higher than ethanol's and methanol's.

Factor 2: External Pressure

Higher external pressure raises boiling point (e.g. pressure cooker); lower pressure lowers it (e.g. at high altitude).

Examples: water boils at 98 deg C at Murree Hills (about 700 torr) and at 69 deg C on Mount Everest (about 323 torr), compared to 100 deg C at sea level (760 torr).

Q11. Define solids and write their general properties.

Solids

Definition: solids are rigid, hard substances with a definite shape and definite volume.

Their constituent atoms, ions, or molecules are closely packed and held together by strong cohesive forces, so the particles cannot move at random.

General Properties of Solids

Very low compressibility and negligible thermal expansion, because particles are already closely packed by strong forces.

Particles only vibrate about fixed mean positions - no translational or rotational motion.

High melting points, resulting from the strong ionic, covalent, or metallic forces holding the particles together.

Q12. Explain the properties of crystalline solids in detail.

(i) Geometrical Shape

Crystalline solids have a definite, distinctive geometric shape due to the orderly 3D arrangement of atoms, ions or molecules.

Interfacial angles remain the same regardless of how the crystal is grown, and faces/angles stay characteristic even when ground to a fine powder.

(ii) Melting Points

Crystalline solids have sharp, definite melting points, which can be used to identify them.

(iii) Cleavage Planes

When broken, crystalline solids split along definite planes called cleavage planes, inclined at a particular characteristic angle for each substance.

(iv) Growing of a Crystal

Crystals grow from a saturated or supersaturated solution by slow evaporation or seeding.

Example: sodium thiosulphate ($Na_2S_2O_3$) solubility is about 231 g/100 cm^3 at 100 deg C versus about 50 g/100 cm^3 at room temperature.

(v) Habit of a Crystal

The shape in which a crystal usually grows is called its habit; this can change under different growth conditions.

Example: NaCl normally forms cubic crystals, but becomes needle-like when 10% urea is present as an impurity.

Q13. What are crystalline and amorphous solids? Give their differences.

Crystalline Solids

Definition: solids with a definite, regular, three-dimensional geometric shape.

Examples: diamond, sodium chloride, ice.

Crystal Lattice: the regular arrangement of ions, atoms, or molecules in three-dimensional space.

Amorphous Solids

Definition: solids that do not possess a regular three-dimensional geometrical shape.

Examples: glass, wood, plastic sulphur, charcoal, coal, coke.

Properties: contain small regions of order called crystallites; melt over a wide temperature range; can be molded and blown into different shapes; have no definite heat of fusion.

Key Differences

Crystalline solids are anisotropic (properties vary with direction) while amorphous solids are isotropic (properties are the same in all directions).

Crystalline solids have sharp melting points; amorphous solids melt gradually over a range.

Quick Check 5.7

(a) Solids have low compressibility and expansion because particles are held very close together by strong forces.

(b) The habit of a crystal is the shape it usually grows into, e.g. NaCl's habit is cubic.

(c) Solids do not undergo translatory motion because of their closely packed arrangement - particles only vibrate about their mean positions.

Q14. Define liquid crystals. Explain their general properties and uses.

Liquid Crystals

Definition: the liquid crystalline state exists between two temperatures - the melting temperature and the clearing temperature.

Example: cholesteryl benzoate forms liquid crystals between 145 deg C and 179 deg C.

Some substances exist in a phase that is neither fully liquid nor fully solid - molecules move like a viscous liquid but with restricted range of motion like a solid: Crystalline Solid -> Cloudy Liquid (Liquid Crystal) -> Clear Liquid.

Molecular shape: rigid, rod-like, with length four to eight times greater than diameter; molecules orient with long axes roughly parallel.

General Properties of Liquid Crystals

Parallel ordered arrangement; elongated, rod-like and linear shape; flow like liquids; show viscosity like liquids; somewhat rigid; always anisotropic.

Uses of Liquid Crystals

(i) Diagnostics: cholesteric liquid crystals can detect tumors and infections by sensing temperature differences between tumors and surrounding tissue, aiding early breast-cancer diagnosis.

(ii) Temperature Sensing: used to identify faulty connections on circuit boards and in thermometers.

(iii) Displays: used in LCDs for TVs, computer screens, calculators, watches, cell phones, and laptops.

Quick Check 5.8

(a) Solid-like properties: parallel ordered arrangement, elongated rod-like shape, optical properties, somewhat rigid, anisotropic.

(b) Liquid-like properties: fluidity, viscosity, ability to fill containers.

(c) Their temperature-sensing property makes them useful for detecting faults in electrical circuits and for use in thermometers.

Section 5 | Answer Key & Formula Summary

MCQ Answer Key (Q1-Q77)

Q.No	Ans	Q.No	Ans	Q.No	Ans	Q.No	Ans
1	B	21	C	41	C	61	B
2	C	22	C	42	C	62	B
3	B	23	C	43	B	63	C
4	A	24	C	44	C	64	C
5	A	25	C	45	C	65	A
6	C	26	C	46	B	66	A
7	C	27	D	47	C	67	C
8	C	28	C	48	C	68	C
9	D	29	C	49	B	69	B
10	C	30	A	50	D	70	C
11	C	31	B	51	C	71	C
12	C	32	C	52	C	72	C
13	C	33	B	53	B	73	D
14	B	34	C	54	C	74	B
15	D	35	B	55	C	75	C
16	D	36	C	56	C	76	C
17	C	37	C	57	C	77	C
18	C	38	C	58	C		
19	B	39	D	59	C		
20	B	40	D	60	C		

Important Formulas

Formula	Description
PV = nRT	Ideal Gas Equation
M = mRT / PV	Molar Mass of a Gas from Gas Equation
V is proportional to 1/P	Boyle's Law (constant n, T)
V is proportional to T	Charles's Law (constant n, P)
V is proportional to n	Avogadro's Law (constant P, T)
Delta Hf = 4.6 kJ/mol (water)	Molar Heat of Fusion
Delta Hv = 40.6 kJ/mol (water)	Molar Heat of Vaporization

Key Constants

Constant	Value

Gas Constant (R)	0.0821 atm dm3 mol-1 K-1 or 8.314 J mol-1 K-1
Molar Heat of Fusion of Water	4.6 kJ/mol
Molar Heat of Vaporization of Water	40.6 kJ/mol
Molar Heat of Vaporization of NH3	21.7 kJ/mol
Molar Heat of Vaporization of HCl	15.6 kJ/mol
Boiling Point of Water (Lahore, ~760 torr)	100 deg C
Boiling Point of Water (Murree, ~700 torr)	98 deg C
Boiling Point of Water (Mount Everest, ~323 torr)	69 deg C
Density Increase of Ice on Freezing	Volume expands by 9% (density decreases)

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