CHAPTER

1 States of Matter

Matter is defined as anything that has mass and occupies space. It exists in several different states such as gaseous, liquid, solid, plasma and Bose–Einstein condensates. Each state of matter has unique physical properties. Gases and liquids take the shape of the containers in which they are placed, whereas solids have their own particular shapes. Gases are easily compressed, but liquids and solids do not. Apart from the above-mentioned states, certain molecules lie between the liquid and crystalline states so called *mesophase* or *liquid crystals*. Supercritical fluids are also considered as mesophase having properties intermediate between those of liquids and gases. Generally, as the temperature or pressure increases matter moves to a more active state and one can observe a physical change.

HIGHLIGHTS

Temperature and *pressure* are the two important factors that determine whether a substance exists in gaseous, liquid or solid state.

GASEOUS STATE

A gas is a compressible fluid and has no definite shape or volume, but occupies the entire container. The kinetic energy of a gas is so high that the effect of intermolecular forces is nil, and thus the intermolecular distances are very large. A gas, at temperatures below its critical temperature, is called a *vapour*, and can be liquefied without cooling by compression alone.

Ideal Gas Law

Several laws that are significantly used to describe the behaviour of gases are as follows:

Boyle's law

This law defines the relationship between the volume of a gas (V) and pressure (P) if the temperature and amount of gas are held constant. According to Boyle's law, at constant temperature, the volume of a gas is inversely proportional to pressure. The law is expressed mathematically as:

$$V \propto \frac{1}{P}$$

or

$$P_1V_1 = P_2V_2 \tag{1.1}$$

Charles' law

This law defines the relationship between volume of a gas and temperature (T) and states that at constant pressure, the volume of a gas is directly proportional to the temperature.

$$\begin{gathered} V \propto T \\ \text{or} \\ \frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad V_1T_2 = V_2T_1 \end{gathered} \tag{1.2}$$

Gay-Lussac's law

This law characterizes the relationship between pressure of the gas and temperature when volume of gas is held constant. According to this law, at constant volume, the pressure of a gas is directly proportional to temperature. The law is expressed mathematically as:

$$\begin{gathered} P \propto T \\ \text{or} \\ \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \text{or} \quad P_1T_2 = P_2T_1 \end{gathered} \tag{1.3}$$

Avogadro's law

Unlike temperature and pressure, volume is an extensive property, which is dependent on the amount of substance present in the system. Avogadro law states that at constant temperature and pressure, the volume of a gas is directly proportional to the number of moles of the gas (n).

$$\begin{gathered} V \propto n \\ \text{or} \\ \frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{or} \quad V_1n_2 = V_2n_1 \end{gathered} \tag{1.4}$$

Combining Eqs. (1.1) and (1.2) with Eq. (1.4) gives the ideal gas law. The ideal gas law is a state of a hypothetical ideal situation and relates temperature, pressure and volume of an ideal or a perfect gas. According to ideal gas law:

$$PV = nRT \tag{1.5}$$

where P is pressure, V volume, n the number of moles, T the temperature and R the gas constant or proportionality constant.

The gas constant is determined experimentally by plotting PV against P and extrapolating to zero pressure (see Fig. 1.1). As one mole of an ideal gas occupies 22.414 L at 1 atmospheric pressure and 0°C, therefore, R is calculated as:

$$R = \frac{PV}{nT} = \frac{1\ \text{atm} \times 22.414\ \text{L}}{1\ \text{mole} \times 273.15\ \text{K}} = 0.08206\ \text{L atm/mole K}$$

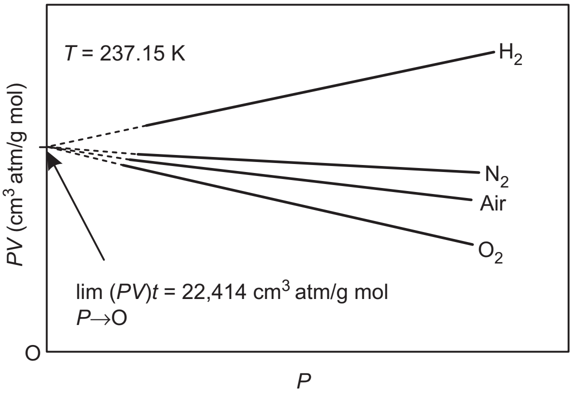

Figure 1.1 Plot of product of pressure and volume of gas versus pressure.

Molecular Weight Determination

If the number of moles of gas (n) is replaced by its equivalent grams of gas (g) per molecular weight (M), then Eq. (1.5) is used to calculate the approximate molecular weight of a gas using equation:

$$PV = \frac{gRT}{M} \tag{1.6}$$

Example 1.1 (Ideal Gas Law)

Calculate the volume occupied by 23.6 g of trifluorochloroethane at 55°C and 720 mmHg pressure.

Solution

According to Eq. (1.6)

$$V = \frac{gRT}{MP}$$

$$= \frac{[(23.6\ \text{g})\ (0.0821\ \text{L atm/mol K})\ (55°\text{C} + 273)]}{[(136\ \text{g/mole})\ (720\ \text{mmHg})(1\ \text{atm}/760\ \text{mmHg})]}$$

$$= 4.9\ \text{L}$$

Real Gases

As the temperature of a gas is lowered and/or its pressure is increased, the ideal gas law is not followed because the volume of the gas is not negligible and intermolecular forces do exist. van der Waals proposed the incorporation of two constant terms, a and b, to account for the deviations from ideal behaviour. The ideal gas law equation then becomes:

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT \tag{1.7}$$

The constant a accounts for the cohesive forces between the gas molecules and constant b accounts for the incompressibility of the gas molecules known as the *excluded volume* occupied by the gas molecules. Due to the cohesive forces between the gas molecules the pressure of the real gas is less than that of an ideal gas. These forces are dependent on the intermolecular distances and related to the density of the gas. The term a/V^2 in Eq. (1.7) is called as *internal pressure per mole*, whereas term $(V - b)$ represents the effective volume of the gas molecules that expand freely. At low pressure conditions, the volume of the gas molecules is large and the contribution of the excluded volume is very small. Under these conditions the term b becomes negligible and Eq. (1.7) is reduced to the ideal gas law (Eq. 1.5).

LIQUID STATE

The liquid state lies between the gaseous and the solid state since there is neither the complete disordered arrangement of constituents as in gases nor the ordered arrangement as in solids. By and large the properties of liquids resemble those of gases while some of the properties approach those of the solids. Like a gas, a liquid can assume the shape of a container and can evenly distribute the applied pressure to every surface in the container. However, unlike a gas, a liquid may not always fill every space in the container, may not compress significantly and will not always mix readily with another liquid. The density of liquid is close to that of solid but unlike a solid, the molecules in a liquid have a much greater freedom to move allowing a liquid to flow.

Properties of liquids can be explained on basis of their following characteristics:

1. Molecules of liquid are in state of random motion but the motion is appreciably smaller in comparison to gases. This explains the incompressibility and higher density of liquids in comparison to gases.
2. The kinetic energy of the molecules of liquid and thus the vapour pressure of the liquid, increases with increase in temperature of the liquid.
3. Force of attraction exists between the molecules of a liquid and is about 10^6 times as strong as in gases. These forces are not strong enough to hold the molecules of liquid in fixed position but are strong enough to disallow them from separating spontaneously. The properties of liquids such as (1) viscosity, (2) surface tension and (3) vapour pressure can be explained in terms of these attractive forces.

Viscosity

Viscosity of a liquid is defined as resistance to flow of liquids. The resistance to flow is developed because of the shearing effect when one layer of liquid moves past another. The detailed description of the viscosity, viscous flow, its measurement and applications has been discussed in Chapter 3.

Surface Tension

A molecule in the bulk of liquid is surrounded by other molecules and is attracted equally in all the directions. The net force on molecule at bulk is zero. However, the molecules on the surface of liquid are subjected to unbalanced forces and are in a higher energy state compared to the bulk phase molecules. The molecules at the surface experience net inward pull because of greater number of molecules per unit volume in the liquid than in the vapour (see Fig. 1.2).

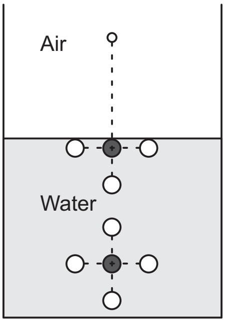

Figure 1.2 Representation of attractive forces on molecules of liquid.

Because of this inward pull, the surface of the liquids tends to contract and attain minimum possible area and behave as if it were in a state of tension. The force that counter balances this inward pull is known as *surface tension*. The detailed description of the surface tension, its measurement and applications has been discussed in Chapter 4.

Vapour Pressure

In case of liquids, kinetic energy is not distributed evenly among molecules and some of the molecules acquire more energy and hence higher velocities than others. The molecules that have sufficient energy to overcome intermolecular attractions are able to escape from the surface into the vapour phase (gas phase). The process is known as *evaporation*. The average kinetic energy of molecules in vapour state is more compared to molecules in liquid state and therefore the temperature of the liquid falls on evaporation. The rate of evaporation of a liquid depends upon the temperature of the liquid, surface area, pressure above the liquid and attractive forces in the liquid. In an another process known as *condensation*, the molecules of liquid in the vapour phase undergo collisions among each other and with the sides of the

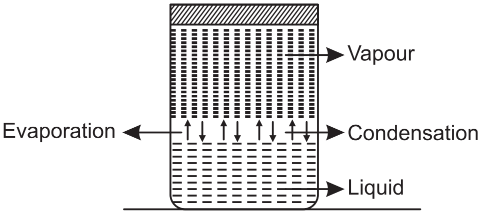

Figure 1.3 Representation of dynamic equilibrium conditions.

evaporating still, transfer their energy to other molecules and come back to the liquid phase. The rate of condensation of molecules in the vapour phase is proportional to the concentration of molecules in the vapour phase. At dynamic equilibrium conditions, the rate of evaporation becomes equal to the rate of condensation, as shown in Figure 1.3.

The relationship between the vapour pressure and the absolute temperature of the liquid is expressed by the Clausius–Clapeyron equation:

$$\log \frac{p_2}{p_1} = \frac{\Delta H\,(T_2 - T_1)}{2.303\,RT_1T_2} \tag{1.8}$$

where, ΔH is the molar heat of vaporization, p_1 and p_2 are the vapour pressure at absolute temperatures T_1 and T_2.

SOLID STATE

In solid state the constituent particles atoms, ions or molecules are packed closely together and have the strongest intermolecular force of attraction. As a result, solids have a stable, definite shape and volume and can only change their shape by force, as when broken or cut. Some characteristic properties of solids are given as follows:

1. Solids are rigid, have definite shape and maintain their volume independent of the container in which they are placed.
2. Solids are nearly incompressible and their compressibility is about 10 times that of gases.
3. Due to closely packed particles, the diffusion of solids is slower compared to liquids or gases.
4. Most solids melt on heating while some undergo sublimation.
5. Solids have high density as compared to gases.

Based on their structural features the solids are classified as follows:

Crystalline Solids

Crystalline solids are those in which the molecules are packed in a definite order, which repeats over and over again throughout the particle. The temperature at which the crystal lattice

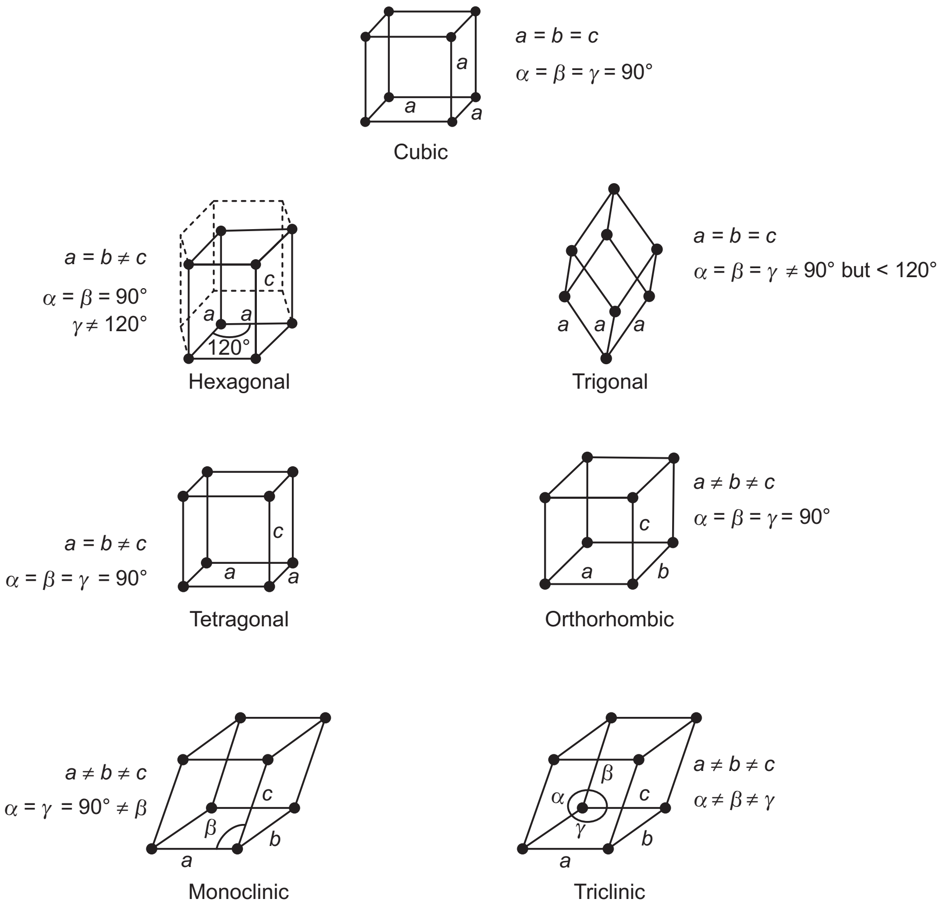

Figure 1.4 Seven possible primitive unit cells.

breaks (by acquiring minimum energy to overcome the withholding attractive forces) is the melting point of the crystal. All crystals are made up of repeating units called *unit cells*. All unit cells in a specific crystal are of the same size and contain the same number of molecules or ions arranged in the same way. Seven primitive unit cells such as cubic, hexagonal, trigonal, tetragonal, orthorhombic, monoclinic and triclinic (see Fig. 1.4) are known. Some of these may also be end centred (monoclinic and orthorhombic), body centred (cubic, tetragonal and orthorhombic) or face centred (cubic and orthorhombic), making a total of 14 possible unit cells called *Bravais lattices* (see Fig. 1.5). Crystalline solids possess a definite and rigid shape. The shape and size of crystals (even of the same materials) differ depending on the conditions under which they are formed. Crystals of a given substance are bound by plane surfaces called *faces*. The angle between any two faces is called as *interfacial angle*. Interfacial angles for a given form will always remain the same and this characteristic is known as *the law of constancy of interfacial angles*.

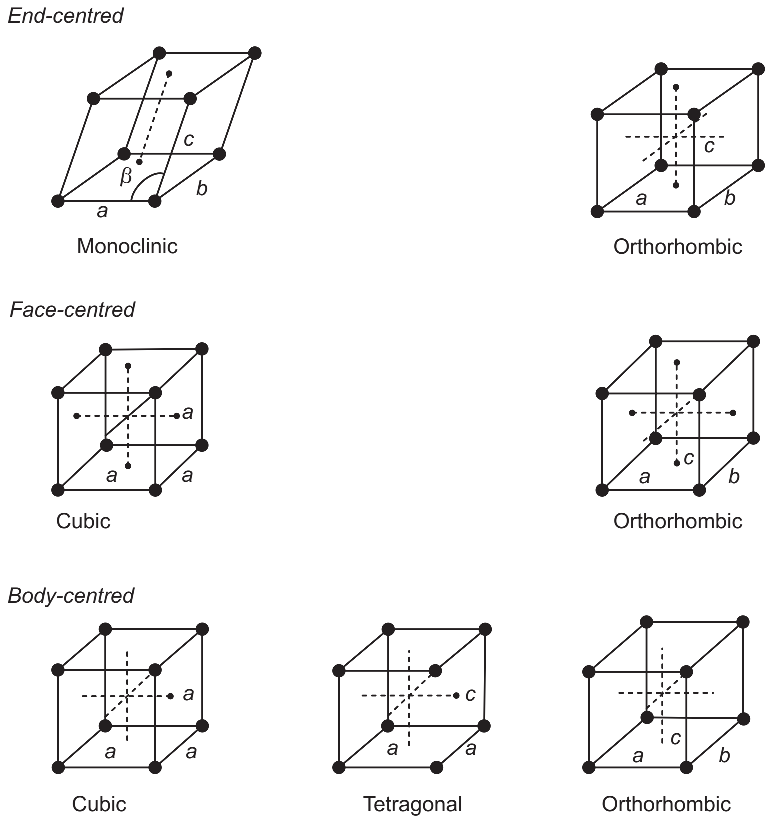

Figure 1.5 End-centred, body-centred and face-centred unit cells.

Some characteristic properties of crystalline solids are as follows:

1. Molecules in crystals are generally held by strong intermolecular forces.
2. Possess characteristic geometrical shapes.
3. They show fracture along a smooth surface when cut or hammered gently.
4. They have sharp melting points.
5. They are anisotropic in nature and their electrical, mechanical and optical properties depend upon the direction along which they are measured.

Types of crystalline solids

Crystalline solids may be classified into the following four types on the basis of the nature of bonds present in them:

Molecular crystals: In molecular crystals, the component particles are molecules held together by weak attraction forces known as *van der Waals forces*. Molecular crystals are soft and compressible, can be distorted very easily, possess low melting and boiling points and are bad conductors of electricity. They are volatile and possess low heats of vaporization and low enthalpy of fusion. Common examples include dry ice, wax, iodine and sulphur.

Ionic crystals: Ionic crystals consist of positively and negatively charged ions arranged in a regular fashion throughout the crystal. They form a three-dimensional network of positive and negative ions in such a way that cations and anions occupy alternate sites. These are held together by strong electrostatic forces. They are very hard and brittle, have very high melting and boiling points and are poor conductors of electricity, but when melted or dissolved in polar solvents, they conduct electricity. Common examples include salts such as sodium chloride and lithium fluoride.

Covalent crystals: In covalent crystals, the constituent particles are atoms of the same or different type, which are bonded to one another by a network of covalent bonds. They are hard and incompressible, extremely nonvolatile and have very high melting points. They are poor conductors of electricity at all temperatures. Common examples include diamond, carborundum (silicon carbide) and quartz (SiO_2).

Metallic crystals: In metallic crystals, the constituent particles are positive kernels, i.e. nuclei where the inner electrons are dispersed in a sea of mobile valence electrons. The forces present between the constituents are metallic bonds. They may be hard as well as soft, are good conductors of heat and electricity, possess metallic lustre, high reflectivity and are highly ductile and malleable, i.e. they can be beaten into sheets and drawn into wires. Common examples include common metals such as nickel, copper and alloys.

Polymorphism

Polymorphism is the ability of a compound to crystallize as more than one distinct crystalline species with different internal lattices. This phenomenon is generally found in any crystalline material including minerals, metals and polymers and is related to allotropy (i.e. the phenomenon of an element existing in two or more physical forms). Polymorphs have different chemical stability and may spontaneously convert from a metastable form to a stable form. Different polymorphic form may have different X-ray diffraction patterns, melting points and solubilities, and these changes affect the drug development program by altering a drug's bioavailability and related parameters. As an example, chloramphenicol palmitate exists in three crystalline polymorphic forms (A, B and C) out of which more soluble form B has higher bioavailability. The formation of polymorph may depend upon several variables pertaining to crystallization process, including the level of supersaturation, temperature of crystallization, geometry of covalent bonds, solvent differences and impurities. The most common example indicating difference in properties of polymorph is the contrast between a graphite and a diamond, both of which are composed of crystalline carbon. Polymorphs can be classified as (1) enantiotropic (one polymorph can be reversibly changed into another by varying temperature or pressure, e.g. sulphur) and/or (2) monotropic (one polymorphic form is unstable at all temperatures and pressures, e.g. glyceryl stearates).

A crystalline solid may contain either a stoichiometric or nonstoichiometric amount of crystallization solvent. Nonstoichiometric adducts, such as inclusions or clathrates, involve entrapped solvent molecules within the crystal lattice. Usually this adduct is undesirable, owing to its lack of reproducibility, and should be avoided for development. A stoichiometric adduct, commonly referred to as a *solvate*, is a molecular complex that has incorporated the crystallizing solvent molecules into specific sites within the crystal lattice. When the incorporated solvent is water, the complex is called a *hydrate*, and the terms hemihydrate, monohydrate and dihydrate describe hydrated forms with molar equivalents of water corresponding to half, one and two. A compound not containing any water within its crystal structure is termed *anhydrous*. During preformulation, it is important to identify the polymorph that is stable at room temperature and to determine whether polymorphic transitions are possible within the temperature range used for stability studies and during processing.

Polycrystalline Solids

In certain crystalline solids, the crystals are very fine and such solids give an impression of being amorphous. Such fine crystalline solid which appears to be amorphous are known as *polycrystalline solids*. Polycrystalline solid appears to be isotropic even though individual crystal is anisotropic.

Amorphous Solids

Amorphous solids possess great disorder and are devoid of any organized structure. In this respect, they resemble liquids. However, their rigidity and cohesiveness allow them to retain a definite shape, and thus for most practical purposes, they can be considered to be solids. Examples of amorphous solids include window glass, polymers such as polystyrene, the silicon in many thin film solar cells and foods such as cotton candy. The lack of molecular order in amorphous solids has a significant effect upon the physical and chemical properties of the sample because the material will have a higher average level of molecular mobility, and a higher entropy and enthalpy than the crystalline form of the same material. The formation of amorphous character within pharmaceutical materials occurs both intentionally (e.g. to improve handling characteristics) and unintentionally (e.g. by poor control of a manufacturing process).

Characteristics of amorphous solids:

1. Amorphous solid soften on heating and gradually begin to flow like liquids.
2. They do not occur in characteristic geometrical shapes.
3. They show fracture in an irregular manner when hammered gently.
4. They do not have sharp melting points.
5. They are isotropic in nature and their electrical, mechanical and optical properties do not depend upon the direction along which they are measured. This property is similar to that of liquids and therefore, amorphous solids are also called as *supercooled liquids*.

LIQUID CRYSTAL STATE

Three states of matter, gas, liquid and solid, have been discussed thus far. A fourth state of matter is the liquid crystal state or mesophase. The liquid crystal state is a distinct state of matter observed between the crystalline solid and liquid states. In the crystalline solid state, the arrangement of molecules is regular (see Fig. 1.6). The molecules are held in fixed positions by intermolecular forces. As the temperature of a substance increases, its molecules vibrate and eventually these vibrations overcome the forces that hold the molecules in place and the molecules start to move. In the liquid state, this motion overcomes the intermolecular forces and the molecules move into random positions (see Fig. 1.6). In the liquid crystal state, the increased molecular motion overcomes the weaker forces, but molecules remain bound by the stronger forces. This produces a molecular arrangement where molecules are in layers, but within each layer, molecules are arranged in random positions, more or less parallel to each other (see Fig. 1.6). The molecules can slide around each other and the layers can slide over one another. This molecular mobility produces the fluidity in liquid crystal state.

Characteristics of Liquid Crystal State

1. The molecules of liquid crystal point along a common axis. This is in contrast to liquid phase molecules, which have no intrinsic order, whereas solid state molecules are highly ordered.
2. Most liquid crystal compounds exhibit polymorphism.
3. They are anisotropic in nature and their properties depend upon the direction along which they are measured.

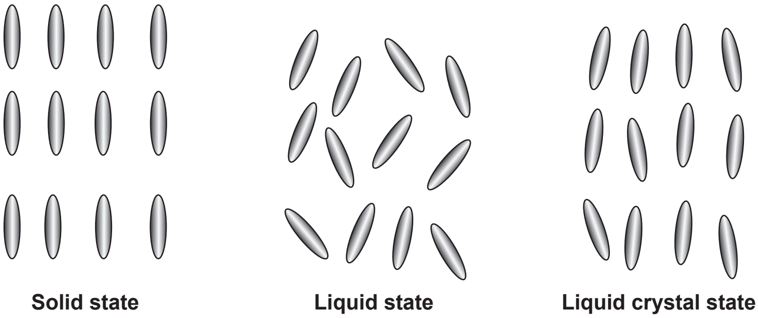

Figure 1.6 Arrangement of molecules in solid, liquid and liquid crystal states.

Types of Liquid Crystals

The three types of liquid crystals are termed *nematic* (thread like), *smectic* (soap like) and *cholesteric*. The nematic state is characterized by molecules that have no positional order but tend to point in the same direction. In the smectic state, the molecules maintain the general orientational order of nematics, but also tend to align themselves in layers or planes. Cholesteric molecules aligned at a slight angle to one another leading to formation of a structure which can be visualized as a stack of very thin two-dimensional nematiclike layers.

SUPERCRITICAL FLUID STATE

A supercritical fluid is a new state of matter where matter is compressible and behaves like a gas (i.e. it fills and takes the shape of its container), but has the typical density of a liquid and hence its characteristics dissolving power. It is a mesophase formed from the gaseous state by application of temperature and pressure that exceeds the critical point of gas. Briefly, the temperature of gas is increased above its critical temperature (liquefaction of gas does not occurs) and then pressure is increased so as to increase the density of gas without significant increase in viscosity (see Fig. 1.7).

The characteristics of superficial fluid state are as follows:

1. It can effuse through solids like a gas, and dissolve materials like a liquid.
2. Does not convert into liquid under pressure change and into gas on increasing the temperature.

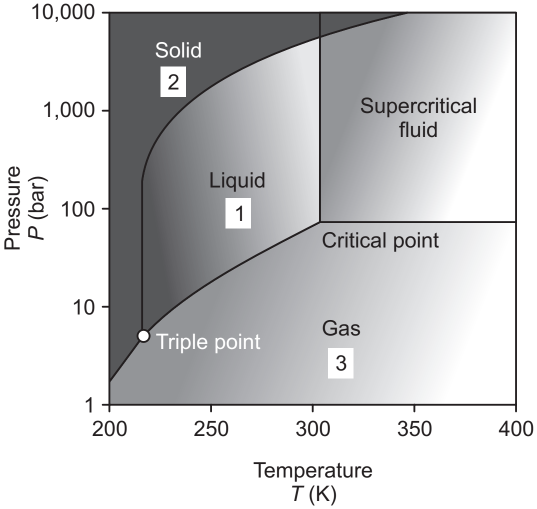

Figure 1.7 Phase diagram of single component system showing supercritical fluid state.

3. Good control over solubility, density, viscosity and other properties of the fluids over a wide range. Close to the critical point, small changes in pressure or temperature result in large changes in density and solubility allowing versatile applications.
4. Absence of surface tension as there is no liquid/gas phase boundary.
5. Carbon dioxide and water are the most commonly used supercritical fluids, being used for decaffeination and power generation, respectively.

There are several important uses of supercritical fluid technology including extraction, chromatography, crystallization and formulation development (sizing of drug substances).

PLASMA

Plasma is an ionized gas to which sufficient energy is provided to free electrons from atoms or molecules and to allow both ions and electrons, to coexist. Plasma is similar to gases but its atoms are made up of free electrons and ions. It is the most abundant form of matter and is formed when enough force is applied on the atomic nucleus to take the electrons. In plasma the interparticle collisions are unlikely and thus plasmas are termed *collisionless*. The example of plasma could be illustrated with the help of fluorescent bulbs having a gas inside the tube. When the light is turned 'on', the electricity flows through the tube, ionizes the gas, excites the atom and creates glowing plasma inside the fluorescent bulb.

HIGHLIGHTS

Bose–Einstein condensate is a gaseous superfluid phase formed by atoms cooled to temperatures very near to absolute zero.

- At ultra-low temperature, a large fraction of the atoms collapse into the lowest quantum state, producing a superfluid.
- It can be thought of as the opposite of plasma.

CHANGES IN THE STATE OF MATTER

Liquefaction of Gases

Any gas can be liquefied by decreasing temperature or increasing pressure. All gas molecules below critical temperature can be liquefied by increasing pressure (see Fig.).

Critical temperature: It is the temperature above which liquefaction of gas does not occurs.

Critical pressure: It is the minimum pressure required to liquefy the gas at its critical temperature.

Critical volume: It is the volume occupied by a mole of gas at critical temperature and pressure.

When temperature of gas is reduced, it loses some of its kinetic energy. If pressure is applied to the gas, the molecules are brought within the sphere of van der Waals interaction forces and thus pass into liquid state. The methods used in liquefaction of gases include:

1. **Faraday's method**: The method is used to liquefy gases whose critical temperature is above or just below atmospheric pressure. The method utilizes freezing mixture to decrease temperature of gases.
2. **Linde's method**: The method is based on the principle that when a compressed gas is allowed to expand into region of low vapour pressure, significant decrease in temperature of gas occurs thus resulting in liquefaction.
3. **Claude's method**: This method is a modification of Linde's method and involves use of cylinder and piston attachment thus part of energy is utilized by the gas in doing mechanical work.

Aerosols

An aerosol is a suspension of fine solid particles or liquid droplets in a gas. The word aerosol derives from the fact that matter 'floating' in air is a suspension (a mixture in which solid or liquid or both solid–liquid particles are suspended in a fluid). In order to differentiate suspensions from true solutions, the term *sol* evolved. With studies of dispersions in air, the term *aerosol* evolved and now embraces liquid droplets, solid particles and combinations of these. The aerosol system depends on the power of compressed or liquefied gas to expel the contents from the container. By pressing the valve excess pressure is created inside container (1–6 atm) that expels the content of the container. As soon as contents are exposed to atmospheric pressure, they get evaporated and form a fine spray. The pressure inside an aerosol container can be achieved by varying the proportions of propellants such as butane, propane, chlorofluorocarbons, nitrous oxide, etc. The advantages of aerosol include the following:

1. Aseptic removal of contents.
2. Direct delivery of medicament to the affected area.
3. No need for mechanical application.
4. Enhanced hydrolytic or oxidative stability.

Boiling Point

Boiling point is the temperature of the liquid at which, the vapour pressure becomes equal to the atmospheric pressure. At this temperature, a liquid changes its state from liquid to vapour. Since at a given pressure different liquids boil at different temperatures, the normal boiling point (also known as the *atmospheric boiling point* or the *atmospheric pressure boiling point*) of a liquid is considered to be taken at an atmospheric pressure at sea level (i.e. 1 atm or 760 mmHg). At the boiling point, all the heat absorbed is used to convert the liquid to the vapour state and there is no increase in the temperature of the liquid until it is completely vapourized. This heat is known as *latent heat of vapourization*.

Melting Point

The melting point of a substance is the temperature at which it changes its state from solid to liquid. During the melting process, all the energy provided to a substance is consumed as latent heat of fusion, and the temperature remains constant. At the melting point, the solid and liquid phases coexist in equilibrium. When considered as the temperature of the reverse change from liquid to solid state, it is referred to as the *freezing point* or *crystallization point*.

Phase Rule

The phase was proposed by J. Willard Gibbs in 1876. It relates number of independent variables or degree of freedom (F), number of phases that can coexist (P) and number of components making up the phases (C) in a system at equilibrium. The least number of independent variables or degree of freedom (i.e. temperature, pressure, concentration, density) can be correlated with number of phases and components for any system at equilibrium using the following equation:

$$F = C - P + 2 \tag{1.9}$$

HIGHLIGHTS

Condensed/reduced phase rule
To reduce complexity in the phase diagram, we consider pressure 1 atm and thus we neglect the existence of vapour phase.
Equation is as follows:

$$F = C - P + 1$$

Single component (C = 1) system (Fig. 1.7)

$$F = C - P + 2$$

$$F = 3 - P$$

At triple point

Three phase ($P = 3$) solid, liquid and vapour coexist, so degree of freedom is

$$F = 3 - P = 3 - 3 = 0$$

This implies that in a system containing single component, this three-phase mixture can only exist at a single temperature and pressure, which is known as a *triple point*.

At points 1, 2 and 3

Single phase exist ($P = 1$), so degree of freedom is

$$F = 3 - P = 3 - 1 = 2$$

This implies that in a single phase condition, two variables—temperature and pressure—can be controlled to any selected pair of values. However, if the single component undergoes a separation into two phases ($P = 2$), F changes from 2 to 1. However, it is not possible to independently control temperature and pressure because change in either of one causes change in other one.

Two-component (C = 2) system containing liquid phases (Fig. 1.8)

Example: Water and phenol partially miscible system

- The miscibility of phenol in water depends on concentration and temperature of the system.
- Curve *gbhci* shows the limits of concentration and temperature within which water and phenol exists in equilibrium.
- Region outside this curve contains one liquid phase.
- Up to 11% phenol concentration (point a to b)—existence of single liquid phase region.
- From 11 to 63% phenol concentration (point b to c)—existence of two-phase region.
- The maximum temperature at which two-phase region exists is termed as *upper consolute temperature* or *critical solution* (66.8°C).
- Line bc is termed as *tie line*. All systems prepared on this line will separate into phases of constant composition called as *conjugate phases*.

Applying phase rule for the water and phenol partially miscible system for the region outside curve ($P = 1$)

$$F = 2 - P + 2$$

$$F = 4 - P = 4 - 1 = 3$$

Considering the system as condensed system, the pressure can be neglected and value of F equalizes 2. Thus, the system is defined by the temperature and concentration of one component.

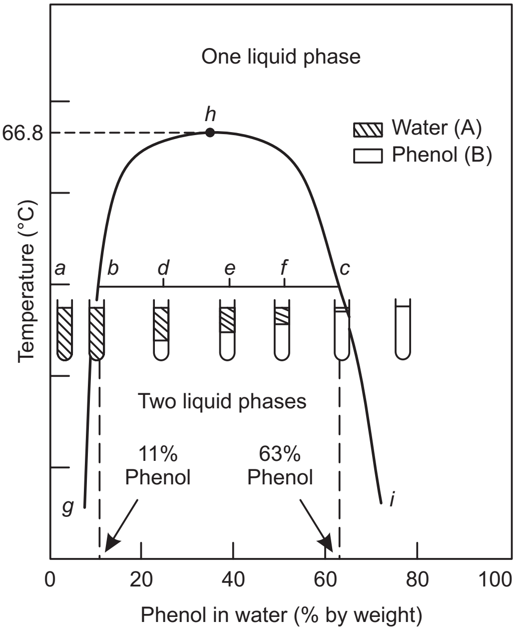

Figure 1.8 Phase diagram for two-component system.

Two-component ($C = 2$) system containing solid and liquid phases (Fig. 1.9)

Example: Eutectic mixtures

The two-components are immiscible as solids but are miscible as liquids.

- Curve AE denotes the freezing of salol and curve BE denotes the freezing of thymol.
- Point A denotes melting point of salol and point B denotes melting point of thymol.
- Region above AEB—existence of single phase system.
- Point E is called as *eutectic point*. At this point, solid salol and thymol and liquid phases of salol and thymol coexist.

Using condensed phase rule:

$$F = C - P + 1 \tag{1.10}$$

At eutectic point E

$$F = 2 - 3 + 1 = 0$$

At any point on curve AE or BE

$$F = 2 - 2 + 1 = 1$$

At point X

$$F = 2 - 1 + 1 = 2$$

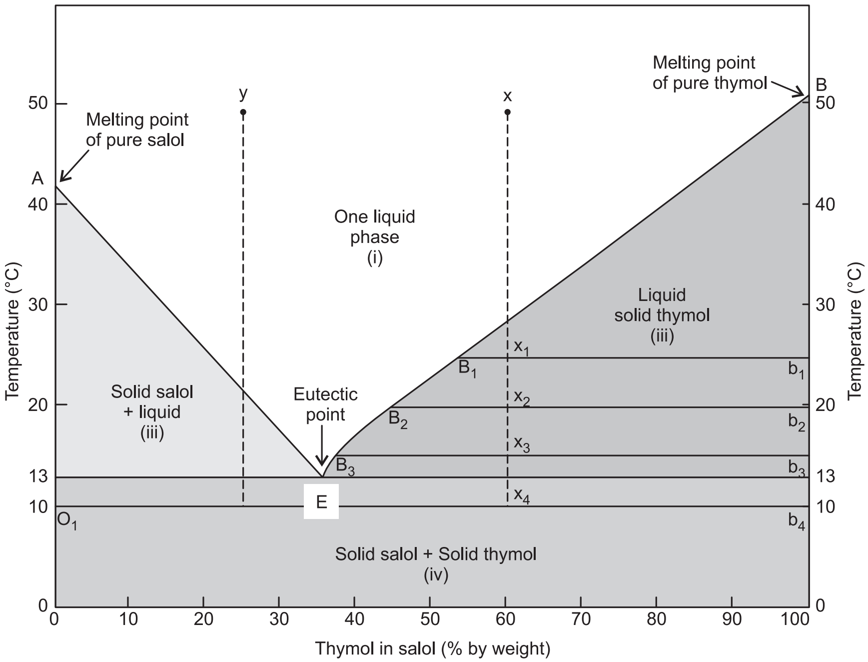

Figure 1.9 Phase diagram for eutectic mixture.

Three-component (C = 3) system (Fig. 1.10)

Applying phase rule for a three-component system, the degree of freedom are as follows:

For single phase

$$F = 3 - 1 + 2 = 4$$

This accounts for temperature, pressure and concentration of any two components.

The phase diagrams for three-component systems are represented by using a triangle whose vertices represents various components present (see Fig. 1.10).

- Area inside triangle represents all the possible combinations of A, B and C to give three-component system.
- At vertex A, B and C—existence of 100% concentration of component A, B and C, respectively.
- From A to B along the line AB—concentration of component B increases.
- From B to C along the line BC—concentration of component C increases.
- From C to A along the line CA—concentration of component A increases.
- At any point inside triangle, the concentration of three components will be added to 100%.
- For finding out concentration of a particular component, we need to draw lines parallel to the base, opposite to vertex.

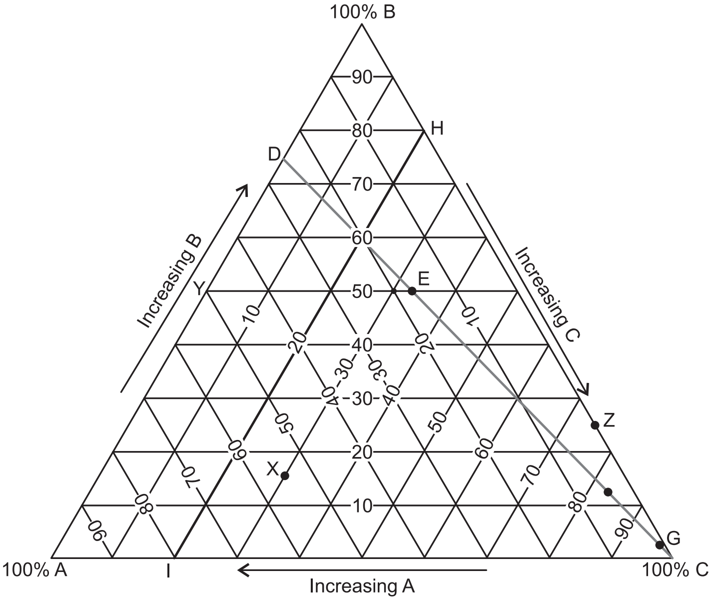

Figure 1.10 Phase diagram for three-component system.

Questions

1. Give proper justification for the following:
   a. At triple point, all the three phases, solid, liquid and gas, exist in equilibrium.
   b. Intermolecular distance between gas molecules is very large.
   c. Volume of a gas is inversely proportional to pressure and directly proportional to the temperature.
   d. Amorphous substances have greater solubility compared with crystalline counterparts.
   e. Weakly basic drugs are better absorbed in the intestine.
2. Write short notes on the following:
   a. Triple point
   b. Gas laws
   c. Type of crystalline solids
   d. Henderson–Hasselbalch equation
   e. Eutectic mixtures
3. Describe in detail the properties of the solid state.
4. Discuss the importance of polymorphism in pharmacy.
5. How does transition occur between states of matter? Describe phase diagram to determine the state of matter at given temperature and pressure.