Physical States of Matter

Major Concepts

Gaseous State

5.1 Typical properties
5.2 Laws related to gases

Liquid State

5.3 Typical Properties

Solid State

5.4 Typical Properties
5.5 Types of Solids
5.6 Allotropy

Time allocation

Teaching periods	16
Assessment periods	04
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Effect on the volume of a gas by a change in the a. pressure b. temperature.
- Compare the physical states of matter with regard to intermolecular forces present between them.
- Account for pressure-volume changes in a gas using Boyle's Law.
- Account for temperature-volume changes in a gas using Charles' Law.
- Explain the properties of gases(diffusion, effusion and pressure).
- Explain the properties of liquids like evaporation, vapour pressure, boiling point.
- Explain the effect of temperature and external pressure on vapour pressure and boiling point.
- Describe the physical properties of solids (melting and boiling points).
- Differentiate between amorphous and crystalline solids. Explain the allotropic forms of solids.

Introduction

Matter exists in three physical states i.e. gas, liquid and solid. The simplest form of matter is the gaseous state. Liquids are less common and most of the matter exists as solid. Matter in gaseous state does not have definite shape and volume. Therefore, gases occupy all the available space. Their intermolecular forces are very weak. Pressure is a significant property of gases. The effect of pressure and temperature on volume of a gas has been studied quite extensively.

The liquid state has strong intermolecular forces hence it has definite volume but it does not have definite shape. It attains the shape of the container in which it is kept. Liquids evaporate and their vapours exert pressure. When vapour pressure of a liquid becomes equal to external pressure, it boils. Liquids are less mobile than gases therefore, they diffuse slowly.

The solid state has definite volume and shape. They are rigid and denser than liquids and gases. They exist in amorphous or crystalline forms.

GASEOUS STATE

5.1 TYPICAL PROPERTIES

Gases have similar physical properties. A few typical properties are discussed here.

5.1.1 Diffusion

Gases can diffuse very rapidly. **Diffusion** *is defined as spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture.* Rate of diffusion depends upon the molecular mass of the gases. Lighter gases diffuse rapidly than heavier ones. For example, H_2 diffuses four times faster than O_2 gas .

5.1.2 Effusion

It is escaping of gas molecules through a tiny hole into a space with lesser pressure. For example, when a tyre gets punctured, air effuses out. Effusion depends upon molecular masses, lighter gases effuse faster than heavier gases.

5.1.3 Pressure

Gas molecules are always in continuous state of motion. Hence, when molecules strike with the walls of the container or any other surface, they exert pressure. **Pressure** ***(P)*** *is defined as the force(F) exerted per unit surface area* ***(A)***.

$$P = F/A$$

The ***SI*** unit of force is **Newton** and that of area is m^2. Hence pressure has ***SI*** unit of $N\,m^{-2}$. It is also called **Pascal** ***(Pa)***

$$\text{One Pascal (Pa)} = 1\ \text{Nm}^{-2}$$

Barometer is used to measure atmospheric pressure and manometer is used to measure pressure in the laboratory.

Standard Atmospheric Pressure

It is the pressure exerted by the atmosphere at the sea level. *It is defined as the pressure exerted by a mercury column of 760 mm height at sea level.* It is sufficient pressure to support a column of mercury 760 mm in height at sea level.

1 atm = 760 mm of Hg = 760 torr (I mm of Hg = one torr)

= 101325 Nm^{-2} = 101325 Pa

5.1.4 Compressibility

Gases are highly compressible due to empty spaces between their molecules. When gases are compressed, the molecules come closer to one another and occupy less volume as compared to the volume in uncompressed state.

5.1.6 Mobility

Gas molecules are always in state of continuous motion. They can move from one place to another because gas molecules possess very high kinetic energy. They move through empty spaces that are available for the molecules to move freely. This mobility or random motion results in mixing up of gas molecules to produce a homogeneous mixture.

5.1.7 Density of Gases

Gases have low density than liquids and solids. It is due to light mass and more volume occupied by the gas molecules. Gas density is expressed in *grams* per dm^3. Whereas, liquid and solid densities are expressed in grams per cm^3 i.e. liquids and solids are 1000 times denser than gases. The density of gases increases by cooling because their volume decreases. For example, at normal atmospheric pressure, the density of oxygen gas is 1.4 g dm^{-3} at 20°C and 1.5 g dm^{-3} at 0°C.

Test yourself 5.1

i. Why the rate of diffusion of gases is rapid than that of liquids?
ii. Why are the gases compressible?
iii. What do you mean by Pascal. How many Pascals are equal to 1 atm?
iv. Why the density of a gas increases on cooling?
v. Why is the density of gas measured in g dm^{-3} while that of a liquid in g cm^{-3}?
vi. Convert the following
a. 70 cm Hg to atm
b. 3.5 atm to torr
c. 1.5 atm to Pa

5.2 LAWS RELATED TO GASES

5.2.1 Boyle's Law

In 1662 **Robert Boyle** studied the relationship between the volume and pressure of a gas at constant temperature. He observed that *volume of a given mass of a gas is inversely proportional to its pressure provided the temperature remains constant.*

According to this law, the volume (V) of a given mass of a gas decreases with the increase of pressure (P) and vice versa. Mathematically, it can be written as:

$$\text{Volume} \propto \frac{1}{\text{Pressure}} \quad \text{or} \quad V \propto \frac{1}{P}$$

$$\text{or} \quad V = \frac{k}{P} \quad \text{or} \quad VP = k = \text{constant}$$

Robert Boyle (1627-1691) was natural philosopher, chemist, physicist and inventor. He is famous for 'Boyle's law of gases'.

Where '***k***' is proportionality constant. The value of ***k*** is same for the same amount of a given gas. Therefore, **Boyle's law** can be *stated as the product of pressure and volume of a fixed mass of a gas is constant at a constant temperature.*

If $P_1V_1 = k$ Then $P_2V_2 = k$

where P_1 = initial pressure P_2 = final pressure

V_1 = initial volume V_2 = final volume

As both equations have same constant therefore, their variables are also equal to each other.

$$\therefore \quad P_1V_1 = P_2V_2$$

This equation establishes the relationship between pressure and volume of the gas.

Experimental Verification of Boyle's law

The relationship between volume and pressure can be verified experimentally by the following series of experiments. Let us take some mass of a gas in a cylinder having a movable piston and observe the effect of increase of pressure on its volume. The phenomenon is represented in figure.5.1. When the pressure of 2 atmosphere (*atm*) is applied, the volume of the gas reads as 1 dm^3. When pressure is increased equivalent to 4 atm, the volume of the gas reduces to 0.5 dm^3. Again when pressure is increased three times i.e. 6 *atm*, the volume reduces to 0.33 dm^3. Similarly, when pressure is increased up to 8 *atm* on the piston, volume of the gas decreases to 0.25 dm^3.

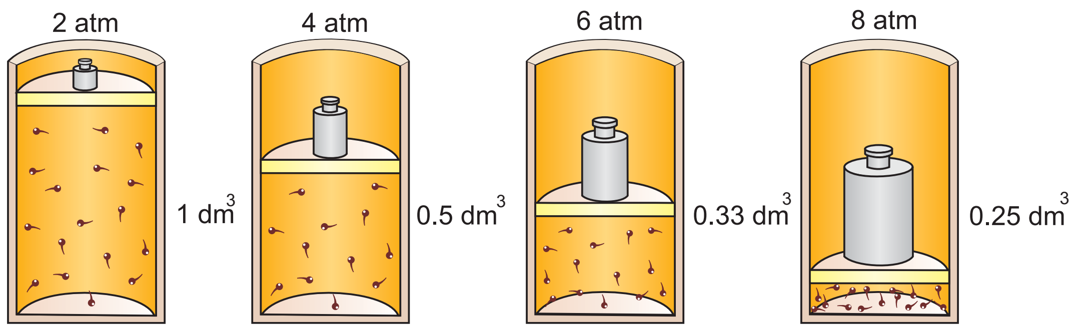


Fig. 5.1 The decrease of volume with increase of pressure.

When we calculate the product of volume and pressure for this experiment, the product of all these experiments is constant i.e. 2 atm dm'. It proves the Boyle's law

$$P_1V_1 = 2\,atm \times 1\ dm^3 = 2\ atm\ dm^3$$
$$P_2V_2 = 4\,atm \times 0.5\ dm^3 = 2\ atm\ dm^3$$
$$P_3V_3 = 6\,atm \times 0.33\ dm^3 = 2\ atm\ dm^3$$
$$P_4V_4 = 8\,atm \times 0.25\ dm^3 = 2\ atm\ dm^3$$

Test yourself 5.2

i. Is the Boyle's law applicable to liquids?

ii. Is the Boyle's law valid at very high temperature?

iii. What will happen if the pressure on a sample of gas is raised three times and its temperature is kept constant?

Do you know?

In which units blood pressure is measured?

Blood pressure is measured using a pressure gauge. It may be a mercury manometer or some other device. Blood pressure is reported by two values, such as 120/80, which is a normal blood pressure. The first measurement shows the maximum pressure when the heart is pumping. It is called ***systolic*** *pressure. When the heart is in resting position, pressure decreases and it is the second value called* ***diastolic****. Both of these pressures are measured in torr units.* ***Hypertension*** *is because of high blood pressure due to tension and worries in daily life. The usual criterion for hypertension is a blood pressure greater than 140/90. Hypertension raises the level of stress on the heart and on the blood vessels. This stress increases the susceptibility of heart attacks and strokes.*

Example 5.1

A gas with volume 350 cm^3 has a pressure of 650 mm of Hg. *If its pressure is reduced to 325 mm of Hg, calculate what will be its new volume*?

Data

$V_1 = 350\ cm^3$

$P_1 = 650\ mm\ of\ Hg$

$P_2 = 325\ mm\ of\ Hg$

$V_2 = ?$

Solution

By using the equation of Boyle's Law

$$P_1V_1 = P_2V_2$$

$$\text{or } V_2 = \frac{P_1V_1}{P_2}$$

By putting the values

$$V_2 = \frac{650 \times 350}{325}$$

$$= 700\ cm^3$$

Thus volume of the gas is doubled by reducing its pressure to half.

Example 5.2

785 cm^3 of a gas was enclosed in a container under a pressure of 600 mm Hg. If volumes is reduced to 350 cm^3, what will be the pressure?

Date

$V_1 = 785\ cm^3$

$P_1 = 600\ mm\ of\ Hg$

$V_2 = 350\ cm^3$

$P_2 = ?$

Solution

By using the Boyle's equation or

$$P_1V_1 = P_2V_2$$

$$or \quad P_2 = \frac{P_1V_1}{V_2}$$

By putting the values or

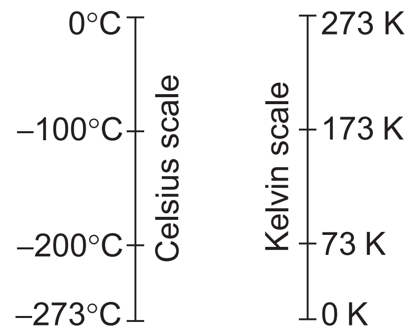
$$P_2 = \frac{785 \times 600}{350} = 1345.7\ mm\ of\ Hg$$

$$or \quad P_2 = \frac{1345.7}{760} = 1.77\ atm$$

Thus pressure is increased by decreasing volume.

Absolute Temperature Scale

Lord **Kelvin** introduced **absolute temperature scale** or **Kelvin scale**. This scale of temperature starts from 0 K or –273.15 °C, which is given the name of **absolute zero**. *It is the temperature at which an ideal gas would have zero volume.* As both scales have equal degree range, therefore, when 0 K equal to –273 °C then 273 K is equal to 0 °C as shown in the scales.



Conversion of Kelvin temperature to Celsius temperature and vice versa can be carried out as follows:

$$(T)\ K = (T)\ °C + 273$$
$$(T)\ °C = (T)\ K - 273$$

5.2.2 Charles's Law

The relationship between volume and temperature keeping the pressure constant was also studied. French scientist **J. Charles** in 1787 presented his law that states *"the volume of a given mass of a gas is directly proportional to the absolute temperature if the pressure is kept constant*'. When pressure ***P*** is constant, the volume ***V*** of a given mass of a gas is proportional to absolute temperature ***T***. Mathematically, it is represented as:

Volume $\propto$ temperature represented as $V \propto T$

or $V = kT$ or $\frac{V}{T} = k$

Where ***k*** is proportionality constant. If temperature of the gas is increased, its volume also increases. When temperature is changed from T_1 to T_2, the volume changes from V_1 to V_2. The mathematical form of Charles' Law will be:

$$V_1 / T_1 = k \quad \text{and} \quad V_2 / T_2 = k$$

As both equations have same value of constant, therefore, their variables are also equal to each other

$$\therefore \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Experimental Verification of Charles's Law

Let us take a certain amount of gas enclosed in a cylinder having a movable piston. If the initial volume of the gas V_1 is 50 cm^3 and initial temperature T_1 is 25 °C, on heating the cylinder up to 100 °C, its new volume V_2 is about 62.5 cm^3. The increase in temperature, increases the volume that can be observed as elaborated below in the figure 5.2.

J. Charles (1746-1823) was a French inventor, scientist, mathematician and balloonist. He described in 1802, how gases tend to expand

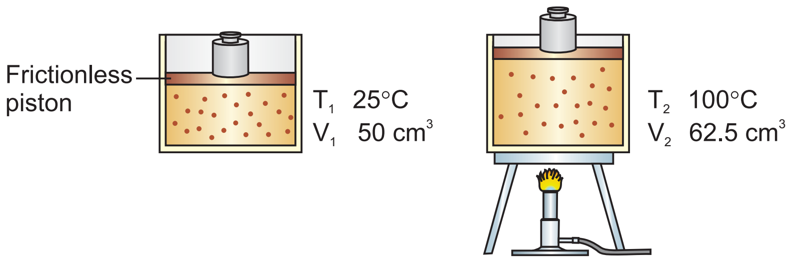


Fig. 5.2: Representation of increase of volume with the increase in temperature.

Remember

- Always convert temperature scale from °C to K scale while solving problems. K = 273 + °C

Example 5.3

A sample of oxygen gas has a volume of 250 cm^3 at –30 °C. If gas is allowed to expand up to 700 cm^3 at constant pressure, find out its final temperature.

Data

$V_1 = 250\ cm^3$

$T_1 = -30\ °C = (-30 + 273) = 243\ K$

$V_2 = 700\ cm^3$

$T_2 = ?$

Solution

By using the equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$or \quad T_2 = \frac{V_2 T_1}{V_1}$$

By putting the value in equation

$$T_2 = \frac{700 \times 243}{250} = 680.4\ K$$

Thus expansion is caused due to increasing temperature

Example 5.4

A sample of hydrogen gas occupies a volume 160 cm^3 at 30 °C. If its temperature is raised to 100 °C, calculate what will be its volume if the pressure remains constant.

Data

V_1	=	160 cm^3		
T_1	=	30 °C =	*303* K	(as 0 °C = *273* K)
T_2	=	100 °C =	*373* K	
V_2	=	?		

Solution

By using the equation of Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

or

$$V_2 = \frac{V_1 T_2}{T_1}$$

By putting the values in above equation.

$$V_2 = \frac{160 \times 373}{303} = 196.9 \ cm^3$$

Thus volume of the gas has increased by raising the temperature.

Remember!

- Degree sign (°) is used with Celsius scale not with Kelvin scale.

Test yourself 5.3

i. Which variables are kept constant in Charles's law?
ii. Why volume of a gas decreases with increase of pressure?
iii. What is absolute zero?
iv. Does Kelvin scale show a negative temperature?
v. When a gas is allowed to expand, what will be its effect on its temperature?
vi. Can you cool a gas by increasing its volume?

Do you know?

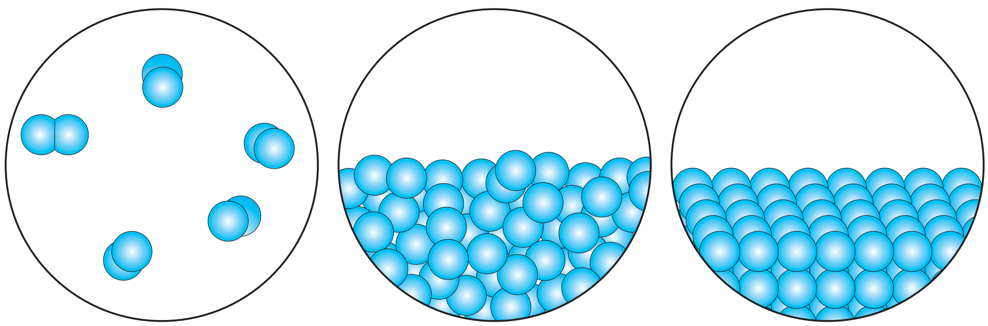
In which units' body temperature is measured?

Body temperature is measured in Fahrenheit scales. Normal body temperature is 98.6 °F, it is equivalent to 37 °C. This temperature is close to average normal atmospheric temperature. In winter atmospheric temperature falls lower than that of our body temperature.
According to principle of heat flow, heat flows out from our body and we feel cold. To control this outward flow of heat, we wear black and warm clothes. To maintain body temperature we use dry fruits, tea, coffee and meats, etc.

Physical States of Matter and Role of Intermolecular Forces

As you know that matter exists in three physical states; gas, liquid and solid. In the gaseous state, the molecules are far apart from each other. Therefore, intermolecular forces are very weak in them. But in the liquid and solid states intermolecular forces play a very important role on their properties.

In the liquid state molecules are much closer to each other as compared to gases as shown in figure 5.3. As a result liquid molecules develop stronger intermolecular forces, which affect their physical properties like diffusion, evaporation, vapour pressure and boiling point. Compounds having stronger intermolecular forces have higher boiling points, as you will see in section 5.3.3.



A. Gas molecules B. Liquids molecules C. Solid molecules

Fig 5.3 Three states of matter showing intermolecular forces.

The intermolecular forces become so dominant in solid state that the molecules look motionless. They arrange in a regular pattern therefore they are denser than molecules of liquids.

LIQUID STATE

Liquids have a definite volume but their shape is not definite. A liquid attains shape of the container in which it is put. A few typical properties of the liquids are discussed here.

5.3 TYPICAL PROPERTIES

5.3.1 Evaporation

The process of changing of a liquid into a gas phase is called ***evaporation***. It is reverse to condensation in which a gas changes into liquid. Evaporation is an endothermic process (heat is absorbed). Such as when one mole of water in liquid state is converted into vapour form, it requires 40.7 kJ of energy.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)} \qquad \Delta H^\circ_{vap} = 40.7 kJ\ mol^{-1}$$

In the liquid state, molecules are in a continuous state of motion. They possess kinetic energy but all the molecules do not have same kinetic energy. Majority of the molecules have average kinetic energy and a few have more than average kinetic energy.

The molecules having more than average kinetic energy overcome the attractive forces among the molecules and escape from the surface. It is called as evaporation.

Evaporation is a continuous process taking place at all temperatures. The rate of evaporation is directly proportional to temperature. It increases with the increase in temperature because of increase in kinetic energy of the molecules.

Evaporation is a cooling process. When the high kinetic energy molecules vapourize, the temperature of remaining molecules falls down. To compensate this deficiency of energy, the molecules of liquid absorb energy from the surroundings. As a result the temperature of surroundings decreases and we feel cooling. For example, when we put a drop of alcohol on palm, the alcohol evaporates and we feel cooling effect.

Evaporation depends upon following **factors**:

i. ***Surface area***: Evaporation is a surface phenomenon. Greater is surface area, greater is evaporation and vice verse. For example, sometimes a saucer is used if tea is to be cooled quickly. This is because evaporation from the larger surface area of saucer is more than that from the smaller surface area of a tea cup.

ii. ***Temperature***: At high temperature, rate of evaporation is high because at high temperature kinetic energy of the molecules increases so high that they over-come the intermolecular forces and evaporate rapidly. For example, water level in a container with hot water decreases earlier than that of a container with cold water. This is because the hot water evaporates earlier than the cold water.

iii. ***Intermolecular forces***: If intermolecular forces are stronger, molecules face difficulty in evaporation. For example, water has stronger intermolecular forces than alcohol, therefore, alcohol evaporates faster than water.

5.3.2 Vapour Pressure

The pressure exerted by the vapours of a liquid at equilibrium with the liquid at a particular temperature is called ***vapour pressure*** *of a liquid.*

The equilibrium is a state when rate of vapourization and rate of condensation is equal to each other but in opposite directions.

$$\text{Liquid} \quad \underset{\text{condense}}{\overset{\text{vapourize}}{\rightleftharpoons}} \quad \text{Vapours}$$

From the open surface of a liquid, molecules evaporate and mix up with the air but when we close a system, evaporated molecules start gathering over the liquid surface. Initially the vapours condense slowly to return to liquid. After sometime condensation process increases and a stage reaches when the rate of evaporation becomes equal to rate of condensation. At that stage the number of molecules evaporating will be equal to the number of molecules coming back(condensing) to liquid. This state is called **dynamic equilibrium** as shown in figure 5.4.

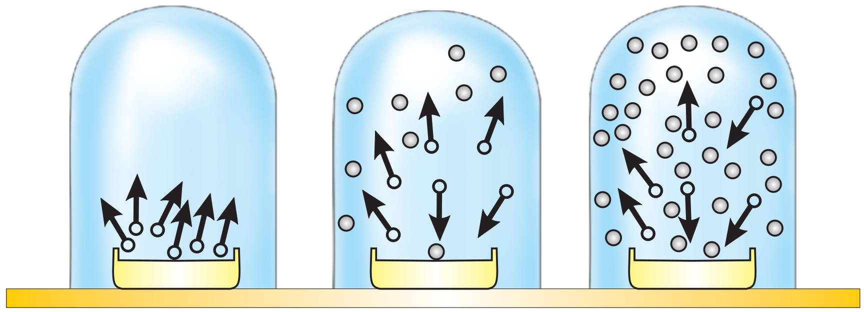


Fig. 5.4 A state of Dynamic Equilibrium between liquid and its vapours

Vapour pressure of a liquid depends upon the following factors.

i. ***Nature of liquid:*** Vapour pressure depends upon the nature of liquid. Polar liquids have low vapour pressure than non-polar liquids at the same temperature. This is because of strong intermolecular forces between the polar molecules of liquids. For example, water has less vapour pressure than that of alcohol at same temperature.

ii. ***Size of molecules:*** Small sized molecules can easily evaporate than big sized molecules hence, small sized molecular liquids exert more vapour pressure. For example, hexane (C_6H_{14}) has a small sized molecule as compared to decane ($C_{10}H_{22}$).

Therefore, C_6H_{14} evaporates rapidly and exerts vapour more pressure than $C_{10}H_{22}$.

iii. ***Temperature:*** At high temperature, vapour pressure is higher than at low temperature. At elevated temperature, the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

For example, vapour pressure of water at different temperatures is given in the Table 5.1.

Table 5.1 Relationship of Vapour Pressure of Water with Temperature

Temp ***°C***	**Vapour Pressure** ***mmHg***	**Temp** ***°C***	**Vapour Pressure** ***mmHg***
0	4.58	60	149.4
20	17.5	80	355.1
40	55.3	100	760.0

5.3.3 Boiling Point

When a liquid is heated, its molecules gain energy. The number of molecules which have more than average kinetic energy increases. More and more molecules become energetic enough to overcome the intermolecular forces. Due to this, rate of evaporation increases that results in increase of vapour pressure until a stage reaches where the vapour pressure of a liquid becomes equal to atmospheric pressure. At this stage, the liquid starts boiling. Hence, **boiling point** *is defined as the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure*.

The figure 5.5 shows the increase of vapour pressure of diethyl ether, ethyl alcohol and water with the increase of temperature. At 0°C the vapour pressure of diethyl ether is 200 mm Hg, of ethyl alcohol 25 mm Hg while that of water is about 5 mm Hg. When they are heated, vapour pressure of diethyl ether increases rapidly and becomes equal to atmospheric pressure at 34.6°C, while vapour pressure of water increases slowly because intermolecular forces of water are stronger. The figure shows the vapour pressure increases very rapidly when the liquids are near to boiling point.

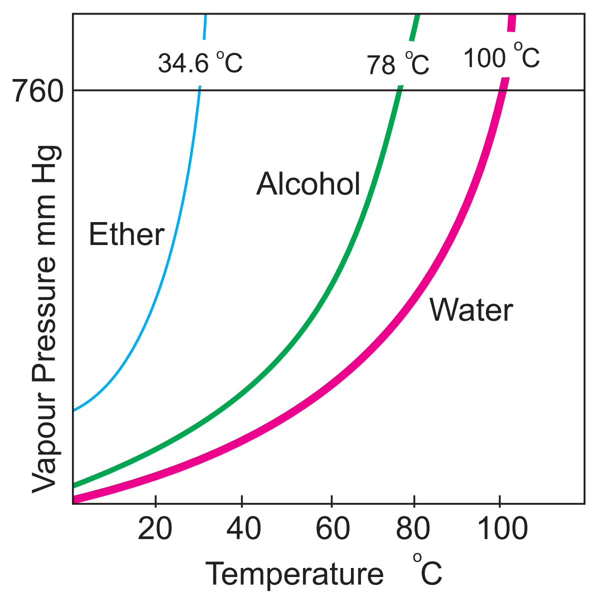


Fig. 5.5 Boiling point curves of Ether, Alcohol and Water.

The boiling point of the liquid depends upon the following factors.

i. ***Nature of liquid:*** The polar liquids have higher boiling points than that of non-polar liquids because polar liquids have strong intermolecular force. Boiling points of a few liquids are given in the table 5.2

ii. ***Intermolecular forces:*** Intermolecular forces play a very important role on the boiling point of liquids. Substances having stronger intermolecular forces have high boiling points, because such liquids attain a level of vapour pressure equal to external pressure at high temperature. It is given in figure 5.5.

iii. ***External pressure:*** Boiling points of a liquid depends upon external pressure. Boiling point of a liquid is controlled by external pressure in such a way, that it can be increased by increasing external pressure and vice versa. This principle is used in the working of 'Pressure Cooker'.

5.3.4 Freezing Point

When liquids are cooled, the vapour pressure of liquid decreases and a stage reaches when vapour pressure of a liquid state becomes equal to the vapour pressure of the solid state. At this temperature, liquid and solid coexist in dynamic equilibrium and this is called the **freezing point** of a liquid. Boiling point and freezing point of a few liquids are given in the table 5.2

Table 5.2 Freezing and Boiling Points of Common Liquids

Sr. No	Liquid	Freezing Point °C	Boiling Point °C
1	Diethyl ether	–116	34.6
2	Ethyl alcohol	–115	78
3	Water	0.0	100
4	n-Octane	–57	126
5	Acetic acid	16.6	118

5.3.5 Diffusion

The liquid molecules are always in a state of continuous motion. They move from higher concentration to lower concentration. They mix up with the molecules of other liquids, so that they form a homogeneous mixture. For example, when a few drops of ink are added in a beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place. Liquids diffuse like gases but the rate of diffusion of liquid is very slow.

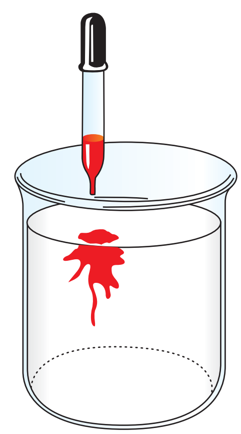


Fig. 5.6 Diffusion in liquids

The diffusion of liquid depends upon the following factors.

i. ***Intermolecular forces:*** Liquids having weak intermolecular forces diffuse faster than those having strong intermolecular forces.

ii. ***Size of molecules:*** Big sized molecules diffuse slowly. For example, honey diffuses slowly in water than that of alcohol in water.

iii. ***Shapes of molecules:*** Regular shaped molecules diffuse faster than irregular shaped molecules because they can easily slip over and move faster.

iv. ***Temperature:*** Diffusion increases by increasing temperature because at high temperature the intermolecular forces become weak due to high kinetic energy of the molecules.

5.3.6 Density

The density of liquid depends upon its mass per unit volume. Liquids are denser than gases because molecules of liquid are closely packed and the spaces between their molecules are negligible. As the liquid molecules have strong intermolecular forces hence they cannot expand freely and have a fixed volume. Like gases, they cannot occupy all the available volume of the container that is the reason why densities of liquids are high. For example: density of water is 1.0 $g\,cm^3$ while that of air is 0.001 $g\,cm^3$. That is the reason why drops of rain fall downward. The densities of liquids also vary. You can observe kerosene oil floats over water while honey settles down in the water.

***Test yourself* 5.4**

I.	*Why does evaporation increase with the increase of temperature?*
ii.	*What do you mean by condensation?*
iii.	*Why is vapour pressure higher at high temperature?*
iv.	*Why is the boiling point of water higher than that of alcohol?*
v.	*What do you mean by dynamic equilibrium ?*
vi.	*Why are the rates of diffusion in liquids slower than that of gases?*
vii.	*Why does rate of diffusion increase with increase of temperature?*
viii.	*Why are the liquids mobile?*

SOLID STATE

It is third state of matter which has definite shape and volume. In solid state, the molecules are very close to one another and they are closely packed. The intermolecular forces are so strong that particles become almost motionless. Hence, they cannot diffuse. Solid particles possess only vibrational motion.

5.4 TYPICAL PROPERTIES

Solids exhibit typical properties, a few of which are discussed here.

5.4.1 Melting point

The solid particles possess only vibrational kinetic energy. When solids are heated, their vibrational energies increase and particles vibrate at their mean position with a higher speed. If the heat is supplied continuously, a stage reaches at which the particles leave their fixed positions and then become mobile. At this temperature solid **melts**. The temperature at which the solid starts melting and coexists in dynamic equilibrium with liquid state is called ***melting point***. The ionic and covalent solids make network structure to form macromolecules. So all such solids have very high melting points.

$$\text{Solid} \rightleftharpoons \text{Liquid}$$

5.4.2 Rigidity

The particles of solids are not mobile. They have fixed positions. Therefore, solids are rigid in their structure.

5.4.3 Density

Solids are denser than liquids and gases because solid particles are closely packed and do not have empty spaces between their particles. Therefore, they have the highest densities among the three states of matter. For example, density of aluminium is 2.70 $g\ cm^{-3}$, iron is 7.86 $g\ cm^{-3}$ and gold is 19.3 $g\ cm^{-3}$.

5.5 Types of Solids

According to their general appearance solids can be classified into two types: amorphous solids and crystalline solids.

5.5.1 Amorphous Solids

Amorphous means shapeless. *Solids in which the particles are not regularly arranged or their regular shapes are destroyed, are called* ***amorphous solids***. They do not have sharp melting points. Plastic, rubber and even glass are amorphous solids as they do not have any sharp melting points.

5.5.2 Crystalline Solids

Solids in which particles are arranged in a definite three-dimensional pattern are called ***crystalline solids***. They have definite surfaces or faces. Each face has definite angle with the other. They have sharp melting points. Examples of crystalline solids are diamond, sodium chloride, etc.

5.6 Allotropy

The existence of an element in more than one forms in same physical state is called **allotropy**. Allotropy is due to:

i. The existence of two or more kinds of molecules of an element each having different number of atoms such as allotropes of oxygen are oxygen (O_2) and ozone (O_3)

ii. Different arrangement of two or more atoms or molecules in a crystal of the element. Such as, sulphur shows allotropy due to different arrangement of molecules (S_8) in the crystals.

They always show different physical properties but have same chemical properties.

Allotropes of solids have different arrangement of atoms in space at a given temperature. The arrangement of atoms also change with the change of temperature and new allotropic form is produced. *The temperature at which one allotrope changes into another is called* ***transition temperature***. For example, transition temperature of sulphur is 96 °C. Below this temperature rhombic form is stable. If rhombic form is heated above 96 °C, its molecules rearrange themselves to give monoclinic form.

$$S_8\ \text{(rhombic)} \xrightleftharpoons{96\ °C} S_8\ \text{(monoclinic)}$$

Other examples are tin and phosphorus.

$$P_4\ \text{(white)} \xrightleftharpoons{250\ °C} (P_4)_n\ \text{(red)}$$

$$\text{Sn}\ _{grey}\ (\text{cubic}) \overset{13.2°\ C}{\rightleftharpoons} \text{Sn}\ _{white}\ (\text{tetragonal})$$

White phosphorus is very reactive, poisonous and waxy solid. It exists as tetra-atomic molecules. While red phosphorous is less reactive, non-poisonous and a brittle powder.

Test yourself 5.5

i. Which form of sulphur exists at room temperature?
ii. Why is white tin available at room temperature?
iii. Why is the melting point of a solid considered its 'identification' characteristic?
iv. Why amorphous solids do not have sharp melting points while crystalline solids do have?
v. Which is lighter one aluminium or gold?
vi. Write the molecular formula of a sulphur molecule?
vii. Which allotropic form of carbon is stable at room temperature (25 °C)?
viii. State whether allotropy is shown by elements or compounds or both?

Curing with salt to preserve meat

Table salt is the most important ingredient for curing meat and is used in large quantities. Salt kills and inhibits the growth of putrifying bacteria by drawing water out of the meat. Concentrations of salt up to 20% are required to kill most species of unwanted bacteria. Once properly salted, the meat contains enough salt to prevent the growth of many undesirable microbes.

CHANGE OF INSTRUMENTATION AS THE SCIENCE PROGRESSES

There are many aspects to be considered about the functioning of instruments. Scientific observation is determined by the human sensory system. It generally relies on instruments that serve as mediators between the world and the senses. Thus, instruments can be considered as a reinforcement of the senses. They provide a great capacity for increasing the power of observation and making induction processes easier. Furthermore, scientific instruments constitute a major factor in checking, refuting or changing previously established theories.

Key Points

- Gases diffuse very rapidly. Diffusion is mixing up of a gas throughout a space or other gases.
- Effusion is escaping of a gas molecule through a fine hole into an evacuated space.
- Gases exert pressure. The SI unit of pressure is Nm which is also called Pascal.
- Standard atmospheric pressure is the pressure exerted by a mercury column of 760 mm height at sea level, it is equivalent to 1 atmosphere.

- Gases are highly mobile and they can be compressed.
- Gases are 1000 times lighter than liquids or solids hence their density is measured in g dm^3.
- Boyle's law states that volume of a given mass of a gas is inversely proportional to the pressure at constant temperature.
- Charles' Law states that volume of a given mass of a gas is directly proportional to the absolute temperature at a constant pressure.
- Absolute zero is the temperature at which an ideal gas would have zero volume, it is –273.15 °C.
- The conversion of a liquid into vapours at all temperatures is called evaporation. It is a cooling process.
- Evaporation depends upon surface area, temperature and intermolecular forces.
- Vapour pressure of a liquid is defined as the pressure exerted by the vapours when liquid and vapour states are in dynamic equilibrium with each other.
- Boiling point is the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.
- Boiling point depends upon the nature of liquid, intermolecular forces and external pressure.
- Freezing point of a liquid is that temperature at which vapour pressure of liquid phase is equal to the vapour pressure of the solid phase. At this temperature liquid and solid coexist in dynamic equilibrium with one another.
- Melting point of solid is the temperature at which solid when heated melts and coexist in dynamic equilibrium with liquid.
- Solids are rigid and denser than liquids.
- Solids are classified as amorphous and crystalline .
- Amorphous solids are shapeless and do not have sharp melting point.
- Crystalline solids have definite three dimensional pattern of arrangement of particles .They have sharp melting points.
- The existence of a solid in different physical forms is called allotropy.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **How many times liquids are denser than gases?**
(a) 100 times (b) 1000 times
(c) 10,000 times (d) 100,000 times

2. **Gases are the lightest form of matter and their densities are expressed in terms of:**
(a) mg cm^{-3} (b) g cm^{-3} (c) g dm^{-3} (d) kg dm^{-3}

3. **At freezing point which one of the following coexists in dynamic equilibrium:**
(a) gas and solid (b) liquid and gas
(c) liquid and solid (d) all of these

4. **Solid particles possess which one of the following motions?**
(a) rotational motions (b) vibrational motions
(c) translational motions (d) both translational and vibrational motions

5. **Which one of the following is not amorphous?**
(a) rubber (b) plastic (c) glass (d) glucose.

6. **One atmospheric pressure is equal to how many Pascals:**
(a) 101325 (b) 10325 (c) 106075 * (d) 10523

7. **In the evaporation process, liquid molecules which leave the surface of the liquid have:**
(a) very low energy (b) moderate energy
(c) very high energy (d) none of these

8. **Which one of the following gas diffuses fastest?**
(a) hydrogen (b) helium
(c) fluorine (d) chlorine

9. **Which one of the following does not affect the boiling point?**
(a) intermolecular forces (b) external pressure
(c) nature of liquid (d) initial temperature of liquid

10. **Density of a gas increases, when its:**
(a) temperature is increased (b) pressure is increased
(c) volume is kept constant (d) none of these

11. **The vapour pressure of a liquid increases with the:**
(a) increase of pressure
(b) increase of temperature
(c) increase of intermolecular forces
(d) increase of polarity of molecules

Short answer questions.

1. What is diffusion, explain with an example?
2. Define standard atmospheric pressure. What are its units? How it is related to Pascal?
3. Why are the densities of gases lower than that of liquids?
4. What do you mean by evaporation how it is affected by surface area.
5. Define the term allotropy with examples.
6. In which form sulphur exists at 100 °C.
7. What is the relationship between evaporation and boiling point of a liquid?

Long Answer Questions

1. Define Boyle's law and verify it with an example.
2. Define and explain Charles' law of gases.
3. What is vapour pressure and how it is affected by intermolecular forces.
4. Define boiling point and also explain, how it is affected by different factors.
5. Describe the phenomenon of diffusion in liquids along with factors which influence it.
6. Differentiate between crystalline and amorphous solids.

Numerical

1. Convert the following units:

   (a) 850 mm Hg to atm (b) 205000 Pa to atm

   (c) 560 torr to cm Hg (d) 1.25 atm to Pa
2. Convert the following units:

   (a) 750 °C to K (b) 150 °C to K

   (c)100Kto°C (d)172Kto°C.
3. A gas at pressure 912 mm of Hg has volume 450cm^3. What will be its volume at 0.4 atm.
4. A gas occupies a volume of 800 cm^3 at 1 atm, when it is allowed to expand up to 1200 cm^3 what will be its pressure in mm of Hg.
5. It is desired to increase the volume of a fixed amount of gas from 87.5 to 118 cm^3 while holding the pressure constant. What would be the final temperature if the initial temperature is 23 °C.
6. A sample of gas is cooled at constant pressure from 30 °C to 10 °C. Comment:

   a. Will the volume of the gas decrease to one third of its original volume?

   b. If not, then by what ratio will the volume decrease?

7. A balloon that contains 1.6 dm^3 of air at standard temperature (0 °C) and (latm) pressure is taken under water to a depth at which its pressure increases to 3.0 atm. Suppose that temperature remain unchanged, what would be the new volume of the balloon. Does it contract or expand?

8. A sample of neon gas occupies a volume of 75.0 cm^3 at very low pressure of 0.4 atm. Assuming temperature remain constant what would be the volume at 1.0 atm. pressure?

9. A gas occupies a volume of 35.0 dm^3 at 17 °C. If the gas temperature rises to 34°C at constant pressure, would you expect the volume to double? If not calculate the new volume.

10. The largest moon of Saturn, is Titan. It has atmospheric pressure of 1.6×10^5 Pa. What is the atmospheric pressure in atm? Is it higher than earth's atmospheric pressure?