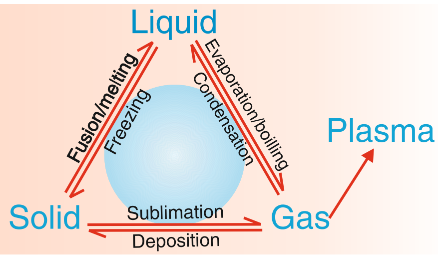
UNIT 5



Physical States of Matter

Unit Outcomes

After completing this unit, you will be able to:

- *understand the kinetic molecular theory and properties of the three physical states of matter;*
- *know the behavior of gases by using the variables volume, temperature, pressure and number of moles;*
- *know terms like ideal gas, diffusion, evaporation, boiling, condensation, vapor pressure, boiling point, molar heat of vaporization, molar heat of condensation, melting, fusion, sublimation, melting point, freezing point, molar heat of fusion, molar heat of solidification;*
- *understand gas laws;*
- *develop skills in solving problems to which the gas laws apply;*
- *perform activities to illustrate gas laws;*
- *carry out experiments to determine the boiling points of liquids and the melting point of solids;*
- *demonstrate an experiment to show phase changes;and*
- *demonstrate scientific inquiry skills: observing, predicting, comparing and contrasting, measuring, interpreting data, drawing conclusion, applying concepts and making generalizations.*

MAIN CONTENTS

Start up Activity

Objective of the activity

Scientists have conducted different activities and experiments to describe the different physical states of matter. The present simple activity will help you to form an idea about the existence of different states of matter (solid, liquid and gas) by using ice (*solid form of water*).

Procedure

1. Take a few pieces of ice in an evaporating dish. Ensure that the ice is present in solid form.
2. Heat the evaporating dish gently and observe the changes.
3. Cover the evaporating dish with a watch glass and continue heating it for more time. Record your observations.

Analysis

1. What happens to ice when it is heated?
2. When the ice is heated in the evaporating dish covered with watch glass, some droplets of water appear on it. What do you conclude from this observation? Can you name this phenomenon?
3. Name the different forms of water that you notice in the above activity.
4. Form a conclusion about the ice as it exists in solid, liquid, and gaseous state. Record your observations and present to the class.

5.1 INTRODUCTION

Competencies

***By the end of this section, you will be able to*:**

- Name the three physical states of matter.

Activity 5.1

Discuss each of the following in your group and present your discussion to the class.

1. Name two examples for each of solids, liquids and gases.
2. What happens when you heat an ice cube?

You recall that all object around us is called matter. Matter is defined as anything that occupies space and has mass. It can exist in the form of gas, liquid and solid. The simplest example is the water we use in our daily life. The three physical states of water are:

- Steam, water in the form of gas.
- Water, in the form of liquid.
- Ice, water in the form of solid.

The changes of the states of matter are our every day experience. That is, ice melts and water freezes; water boils and steam condenses.

The physical state of a given sample of matter depends on the temperature and pressure. Changing these conditions or variables may change the behaviour of the substances as solids, liquids or gases.

Solid

A solid has a definite shape and a definite volume. Solids are almost completely incompressible and have very high average density. A high average density reflects the fact that the particles within solids are usually packed closer than those in liquids or gases do. The tightly packed particles of solids are also highly organized.

The particles of a solid, whether they are atoms, ions or molecules only vibrate about a fixed point with respect to the neighboring particles. Because of these, the particles maintain a fixed position; for example substances like metals, wood, coal and stone, are solids.

Liquid

A liquid has a definite volume, but does not have a definite shape. Liquids take the shape of their container. This is explained in terms of arrangement of particles. In the liquid state, particles vibrate about a point, and constantly shift their positions relative to their neighbours. At room temperature, water, ethanol, benzene and oil are liquids.

Gas

A gas has neither a definite shape nor a definite volume. This is because its particles are virtually independent of one another. For example, air, hydrogen, oxygen, carbon dioxide and nitrogen are gases.

Plasma

Besides, solid, liquid and gas, there exists a fourth-state of matter at very high temperature (million degrees Celsius). At such high temperatures molecules cannot exist. Most or all of the atoms are stripped of their electrons. This state of matter, a gaseous mixture of positive ions and electrons, is called plasma. Because of the extreme temperatures needed for fusion, no material can exist in the plasma state.

Exercise 5.1

1. Can oxygen exist as a liquid and solid?
2. Compare and contrast the three states of matter.
3. What is dry ice?

5.2 KINETIC THEORY AND PROPERTIES OF MATTER

Competencies

***By the end of this section, you will be able to*:**

- give examples for each of the three physical states of matter;
- state kinetic theory of matter;
- explain the properties of the three physical states of matter in terms of kinetic theory; and
- compare and contrast the three physical states of matter.

Activity 5.2

Form a group and perform the following task. Present your findings to the class.

1. Select any three different substances; one existing in the solid state, the second in the liquid state and the third in the gaseous state at room temperature. Use the following table to explain the motion, distance and attraction between particles.

	Substances	Motion of particles	Distance between particles	Attraction between particles
1				
2				
3				

2. In which state do the particles possess the highest kinetic energy?

5.2.1 The Kinetic Theory of Matter

The three states of matter in which substances are chemically the same but physically different are explained by the kinetic theory of matter. The kinetic theory of matter gives an explanation of the nature of the motion and the heat energy. According to the kinetic theory of matter, every substance consists of a very large number of very small particles called ions, atoms and molecules. The particles are in a state of continuous and random motion with all possible velocities. The motion of the particles increases with a rise in temperature.

Generally, the kinetic theory of matter is based on the following three assumptions:

1. All matter is composed of particles which are in constant motion.
2. The particles possess kinetic energy and potential energy.
3. The difference between the three states of matter is due to their energy contents and the motion of the particles.

5.2.2 Properties of Matter

Activity 5.3

Form a group and discuss each of the following idea.Present your discussion to the class.

1. Compare and contrast the density of solid, liquid and gaseous forms of a substance.

2. Discuss the compressibility of solid, liquid and gaseous forms of water.

As discussed earlier, matter exists as gas, liquid, and solid. Their properties are explained in terms of the kinetic theory as follows:

Properties of Gases

From the kinetic molecular theory of gases, the following general properties of gases can be summarized.

1. Gases have no definite shape and definite volume. This is because gases assume the volume and shape of their containers.
2. Gases can be easily compressed. By applying pressure to the walls of a flexible container, gases can be compressed; the compression results in a decrease in volume. This happens due to the large spaces between the particles of gases.
3. Gases have low densities compared with liquids and solids. This is due to the fact that the particles of a gas are very far apart and the number of molecules per unit volume is very small. A small mass of a gas in a large volume results in a very low density.
4. Gases exert pressure in all directions. Gases that are confined in a container exert pressure on the walls of their container. This pressure is due to collisions between gas molecules and the walls of the container.
5. Gases easily flow and diffuse through one another. A gas moves freely and randomly throughout in a given space.

Properties of Liquids

Liquids can be characterized by the following properties.

1. Liquids have a definite volume, but have no definite shape. They assume the shapes of their container. Lack of a definite shape for liquid substances arises from its low intermolecular forces as compared to that of solids.
2. Liquids have higher densities than gases. Their density is a result of the close arrangement of liquid particles. Thus, the particles of liquids are closer than those of gases. This accounts for the higher densities of liquids as compared to gases.
3. Liquids are slightly compressible. With very little free spaces between their particles liquids resist an applied external force and thus are compressed very slightly.
4. Liquids are fluids. A fluid is a substance that can easily flow. Most liquids naturally flow downhill because of gravity. Because liquids flow readily the molecules of a liquid can mix with each other. They flow much more slowly than gases.

Properties of Solids

1. Solids have a definite shape and a definite volume. This is due to the strong force of attraction that holds the particles of solids together.

2. Solids generally have higher densities than gases and liquids. The particles of solids are very close to each other. There is almost no empty space between the particles of solids. This closeness of particles makes solids to have more particles (mass) per unit volume, and hence solids have high density.
3. Solids are extremely difficult to compress. This is because of the high interparticle forces, and a very short distance between the particles.
4. Solids are not fluids. That is they normally do not flow. This is due to the fact that solid particles are rigidly held in position by strong forces that cause the restricted motion of their particles.

Activity 5.4

Form a group and perform the following activity:

Take a balloon and blow air into it. What happens to the volume of balloon. Explain your observations and findings to the class.

Exercise 5.2

Arrange the three states of matter in order of increasing:

a intermolecular force

b density

c compressibility

d kinetic energy

5.3 THE GASEOUS STATE

Competencies

By the end of this section, you will be able to:

- explain the assumptions of kinetic molecular theory of gases;
- describe the properties of gases using kinetic molecular theory;
- describe the behavior of gases by using the variables *V* (volume), *T* (Temperature) *P* (pressure) and *n* (number of moles);
- state Boyle's law;
- perform an activity to show changes in volume and pressure of gases to illustrates Boyle's law;

- apply Boyle's law in solving problems;
- state Charles' law;
- perform an activity to show changes in volume and temperature of gases to illustrate Charles' law;
- apply Charles' law in solving problems;
- derive combined gas law equation from Boyle's law and Charles' law;
- use the combined gas law to calculate changes in volume, pressure or temperature;
- define an ideal gas;
- derive an ideal gas equation from Boyle's law, Charles' law and Avogadro's law;
- compare the nature of real gases with ideal gases;
- solve problems related to ideal gas equation;
- define diffusion;
- state Graham's law of diffusion;
- carry out an activity to compare the rate of diffusion of two different gases;
- apply Graham's law of diffusion in solving problems.

Activity 5.5

Form a group and discuss each of the following phenomenon. Present your conclusion to the class.

Regarding the pressure in the tyres, what would you recommend to the driver of a car who is taking a trip to an area that is experiencing very cold temperatures, more or less air? Explain.

5.3.1 The Kinetic Molecular Theory of Gases

The particles in an ideal gas are very widely spaced and they are in a constant random motion. The pressure of a gas is the result of continuous collisions between the particles and the walls of their container.

Assumptions of the kinetic molecular theory of gases

1. The particles are in a state of constant, continuous, rapid, random motion and, therefore, possess kinetic energy. The motion is constantly interrupted by collisions with molecules or with the container. The pressure of a gas is the effect of these molecular impacts.

2. The volume of the particles is negligible compared to the total volume of the gas. Gases are composed of separate, tiny invisible particles called molecules. Since these molecules are so far apart, the total volume of the molecules is extremely small compared with the total volume of the gas. Therefore, under ordinary conditions, the gas consists chiefly of empty space. This assumption explains why gases are so easily compressed and why they can mix so readily.

3. The attractive forces between the particles are negligible. There are no forces of attraction or repulsion between gas particles. You can think of an ideal gas molecule as behaving like small billiard balls. When they collide, they do not stick together but immediately bounce apart.

4. The average kinetic energy of gas particles depends on the temperature of the gas. At any particular moment, the molecules in a gas have different velocities. The mathematical formula for kinetic energy is $K.E. = \frac{1}{2} mv^2$, where m is mass and v is velocity of gas molecules. Because the molecules have different velocities, they have different kinetic energies. However, it is assumed that the average kinetic energy of the molecules is directly proportional to the absolute (Kelvin) temperature of the gas.

5.3.2 The Gas Laws

Activity 5.6

Form a group and discuss the following phenomena in terms of the gas laws. Present your conclusion to the class.

a The increase in pressure in an automobile tire on a hot day.

b The loud noise heard when a light bulb shatters.

The gas laws are the products of many experiments on the physical properties of gases, which were carried out over hundreds of years. The observation of Boyle and other scientists led to the development of the Gas Laws. The gas laws express mathematical relationships between the volume, temperature, pressure, and quantity of a gas.

Pressure: pressure is defined as the force applied per unit area.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

Pressure is one of the measurable properties of gases. Thus, the pressure of a gas can be expressed in unit of atmosphere, Pascal, torr, millimetre of mercury. The SI unit of pressure is Pascal (Pa), and is defined as one Newton per square metre.

$$1 \text{ Pa} = 1 \text{ N/m}^2 \text{ and}$$

$$1 \text{ atm} = 760 \text{ mmHg} = 76 \text{ cmHg} = 760 \text{ torr} = 101325 \text{ Pa} = 101.325 \text{ kPa}$$

Volume: Volume is the space taken up by a body. The SI unit of volume is the cubic metre (m^3). Volume is also expressed in cubic centimetre (cm^3) and cubic decimetre (dm^3). Other common units of volume are millilitre (mL) and litre (L).

$$1 \text{ cm}^3 = (1\times10^{-2} \text{ m})^3 = 1\times10^{-6} \text{ m}^3$$

$$1 \text{ dm}^3 = (1\times10^{-1} \text{ m})^3 = 1\times10^{-3} \text{ m}^3 = 1 \text{ L}$$

A litre is equivalent to one cubic decimeter: The relation is given as follows

$$1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$$

Temperature: Temperature is the degree of hotness or coldness of a body. Three temperature scales are commonly used. These are °F (degree Fahrenheit), °C (degree Celsius) and K (Kelvin). In all gas calculations, we use the Kelvin scale of temperature.

We use the following formulae for all necessary inter-conversions:

$$K = °C + 273$$

$$°C = (°F - 32)\frac{5}{9}$$

$$°F = \left(\frac{9}{5} \times °C\right) + 32$$

Exercise 5.3

Convert the following:

a 500 mmHg into atm, torr, and cmHg

b 100 dm^3 into mL, cm^3, L, m^3

c 54°C into K and °F.

Molar Volume and Standard Conditions (STP)

The conditions of a pressure of 1 atmosphere and a temperature of 0ºC (273.14 K) are called standard temperature and pressure or STP for gases. At STP the volume of one mole of any gas is equal to 22.4 litres. This volume is known as molar gas volume.

Quantity of gas: The quantity of a gas is expressed in mole (n). Mole is the quantity of gas in terms of number of particles. It is the number of atoms or molecules in 1 gram-atom or 1 gram-molecule of an element or a compound.

1. *Boyle's Law*

Activity 5.7

Form a group and discuss the following phenomenon. Present your discussion to the class.

Explain why a helium weather balloon expands as it rises more in the air. (Assume the temperature remains constant.)

The first quantitative experiments on gases were performed by the Irish chemist, Robert Boyle (1627-1691). His experiment helped to analyze the relationship between the volume and pressure of a fixed amount of a gas at constant temperature. Decreasing the external pressure, causes the gas to expand and to increase in volume. Correspondingly, increasing the external pressure allows the gas to contract and decrease in volume. This is shown in Figure 5.1.

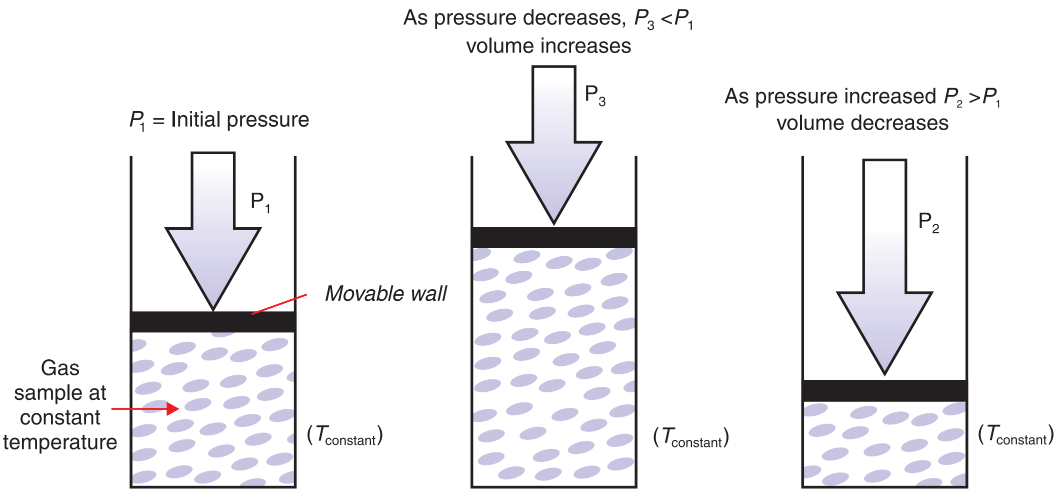


Figure 5.1 The relation between pressure and volume.

Boyle studied the relationship between the pressure of the trapped gas and its volume. Accordingly, he discovered that at constant temperature doubling the pressure on a sample of gas reduces its volume by one-half. Tripling the gas pressure reduces its volume to one-third of the original. Generally, the volume of a gas decreases, as the pressure on the gas increases. This volume-pressure relationship is illustrated in Table 5.1.

Table 5.1 Pressure and volume data for a gas at constant mass and temperature.

Presssure (atm)	Volume (mL)	$P \times V$
0.5	1200	600
1.0	600	600
2.0	300	600
3.0	200	600
4.0	150	600
5.0	120	600
6.0	100	600

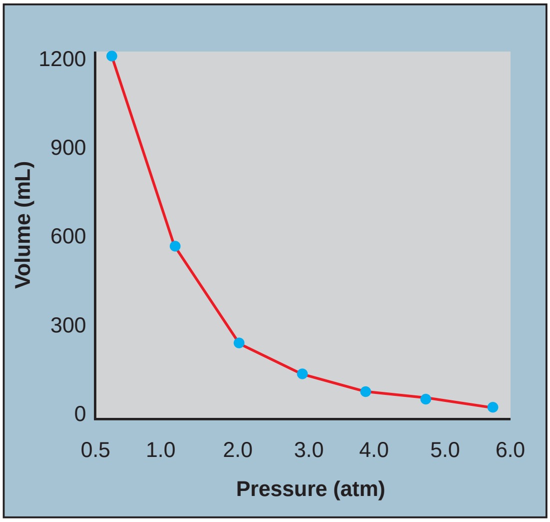


Figure 5.2 Volume versus pressure graph for a gas at constant temperature and mass.

The general volume-pressure relationship illustrated above is called Boyle's law. Boyle's law states that *the volume of a fixed mass of gas is inversely proportional to the pressure at a constant temperature.* The inverse relationship between pressure and volume is mathematically given as

$$V \ \alpha \ \frac{1}{P} \ \text{(at constant } T \text{ and } n\text{)}$$

From which follows,

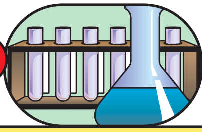
$$V = k\frac{1}{P} \quad \text{or} \quad PV = k;$$

where k is a constant at a specific temperature for a given sample of gas.

If P_1 and V_1 represent the initial conditions; and P_2 and V_2 represent the new or final conditions, Boyle's law can be written as:

$$P_1V_1 = P_2V_2;$$

Experiment 5.1



Effect of Pressure on the Volume of Gas

Objective: To observe the relationships between the volume and pressure of a gas at constant temperature.

Apparatus: U-tubes, ruler, rubber tube, burette, glass tube.

Chemicals: Mercury.

Procedure:

1. Join two tubes by a rubber tubing to give a U-arrangement as shown in Figure 5.3, and then partially fill these two tubes with mercury.
2. Put a ruler in the middle of the tube.
3. The first arm of the tube (A) contains air and is sealed by a tap.
4. By moving the second arm of the tube (B) up and down, the volume of air in the first tube can be varied.
5. The pressure exerted on the air is obtained from the difference in height of mercury in the two arms of the tube.

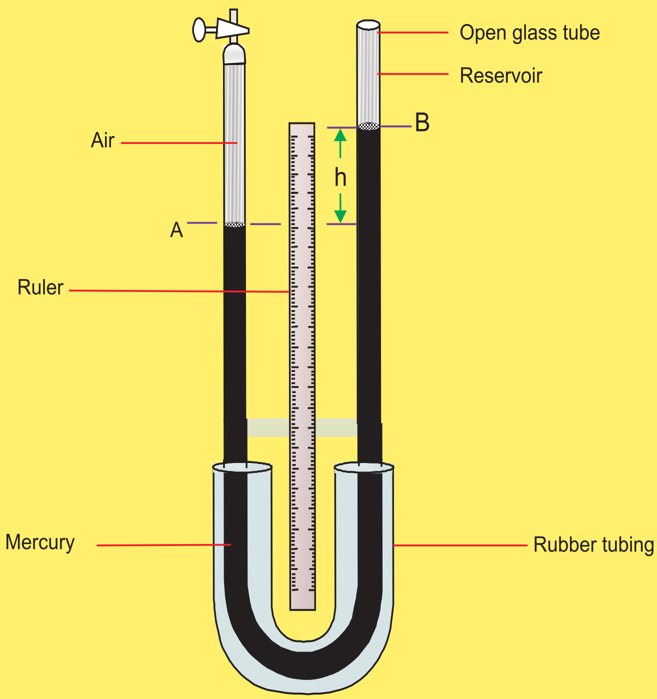


Figure 5.3 Effect of pressure on the volume of a gas at constant temperature.

Observations and analysis:

1. Plot a graph taking pressure on the vertical axis *versus* volume on the horizontal axis and comment on the shape of the graph.
2. What can you conclude from this experiment?

Activity 5.8

Discuss the following activity in your group and present your discussion to the class. Plot a graph of pressure, P, versus 1/V, using the following data:

(Note: Calculate 1/V values from the given volume by yourself)

Pressure(atm)	1	2	3	4	5
Volume(mL)	20	10	6.7	5	4

1. What does the graph look like?
2. What does the graph indicate?
3. What can you conclude from the graph?

Example 1

An inflated balloon has a volume of 0.55 L at sea level (1.0 atm) and is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming that the temperature remains constant, what is the final volume of the balloon?

Solution:

Givens:

Initial conditions	Final conditions
$P_1 = 1.0$ atm	$P_2 = 0.40$ atm
$V_1 = 0.55$ L	$V_2 = ?$

Use Boyles' law equation: $P_1V_1 = P_2V_2$

Therefore, $V_2 = \frac{P_1V_1}{P_2} = 0.55\,\text{L} \times \frac{1.0\,\cancel{\text{atm}}}{0.40\,\cancel{\text{atm}}} = 1.4\,\text{L}.$

Exercise 5.4

1. A certain gas occupies a volume of 10.0 m^3 at a pressure of 100.0 kPa. If its volume is increased to 20 m^3, what would be the new pressure of the gas assuming temperature remains constant?
2. A laboratory experiment shows 4.0 litres of helium gas trapped in a cylinder at a pressure of 7.0 atm. The pressure is decreased, at a constant temperature, to 2.00 atm. What is the new volume?

2. *Charles' Law*

Historical Note

Jacques Alexandre Ceser Charles

The French scientist Charles was most famous in his lifetime for his experiment in ballooning. The first such flights were made in 1783, using a balloon of linen paper filled with hot air. Charles filled a silk balloon with hydrogen. He developed the gas law what we call Charles' law. The volume of samples of gases increases with increasing temperature (at constant pressure).

Activity 5.9

Form a group and discuss the following activity. Present your discussion to the class.

1. Have you ever wondered what makes a pop corn 'pop'?
2. Explain Charles' law in terms of kinetic molecular theory.

The French physicist, Jacques Charles (1746-1823), was the first person to fill a balloon with hydrogen gas and made the first solo balloon flight.

In 1787, Jacques Charles investigated quantitative relationship between the volume and temperature of a fixed quantity of gas which is held at constant pressure.

This can be related to real life by using an empty balloon. Accordingly, if we take a balloon filled with air and place it on boiling water, the volume of the balloon visibly increases as shown in Figure 5.4(b). If we take the balloon and place it in the freezer compartment of our refrigerator, its volume shrinks drastically as the air inside cools (Figure 5.4(c)).

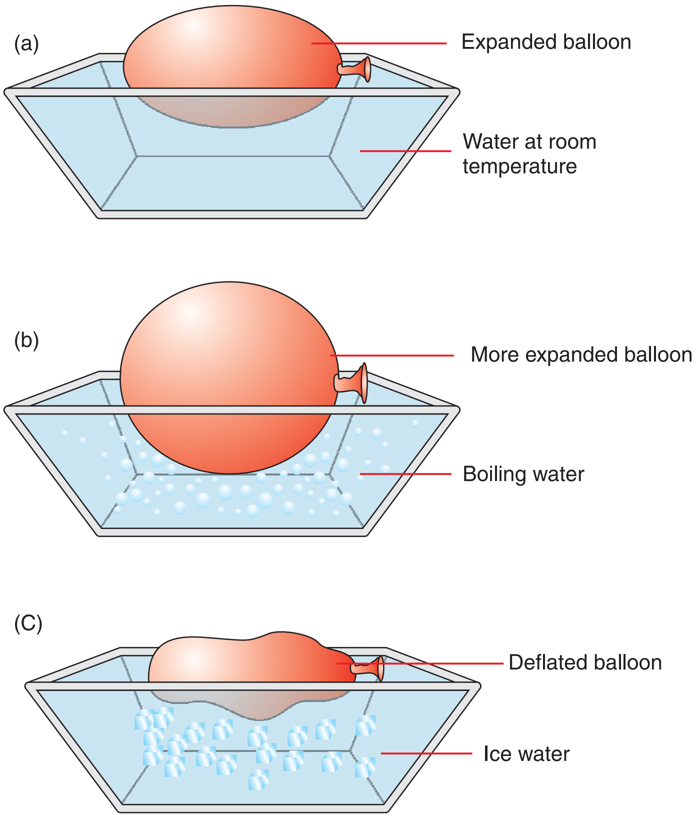


Figure 5.4 Relationship between the volume of air in the balloon and its temperature.

From your observation what can you conclude about the effect of temperature on the volume of a gas?

In 1848, Lord Kelvin realized that a temperature of -273.15°C is considered as absolute zero. Absolute zero is theoretically the lowest attainable temperature. Then he set up an absolute temperature scale, or the Kelvin temperature scale, with absolute zero as the starting point on the Kelvin scale.

The average kinetic energy of gas molecules is closely related to the Kelvin temperature. The volume of a gas and Kelvin temperature are directly proportional to each other. For example, doubling the Kelvin temperature causes the volume of a gas to double, and reducing the Kelvin temperature by half causes the volume of a gas to decrease by half. This relationship between Kelvin temperature and the volume of a gas is known as Charles' law.

Charles' law states that *the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.*

Mathematically;

$$V \propto T \text{ at constant } P \text{ and } n:$$

$$V = kT, \quad \text{or} \quad \frac{V}{T} = k$$

The value of T is the Kelvin temperature, and k is a constant. The value of k depends only on the quantity of gas and the pressure. The ratio V/T for any set of volume-temperature values always equals the same k. Charles' law can be applied directly to volume-temperature problems using the relationship:

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

where V_1 and T_1 represent the initial condition; V_2 and T_2 represent the new condition.

Charles found that the volume of a gas at constant pressure increases linearly with the temperature of the gas. That is, a plot of the volume of a gas versus its temperature (K) gives a straight line assuming the pressure remains constant. Figure 5.5 illustrates the relationship between the volume of a gas and Kelvin temperature. If the graph is extrapolated to give zero volume for a gas, the temperature reached is –273.15°C. This temperature is the lowest temperature attained by a gas and it is called absolute zero.

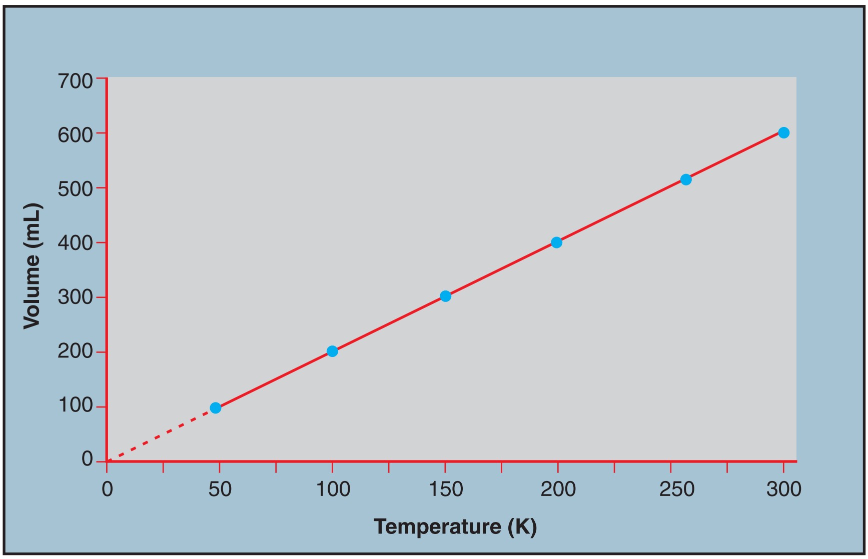


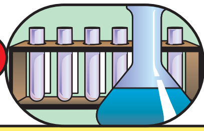
Figure 5.5 A plot of volume versus the Kelvin temperature.

Note that the ratio V/T is constant for every plot of the curve. Figure 5.5 is drawn using the data given in Table 5.2 for the same sample of gas. In all cases, V/T is constant as noted in the table.

Table 5.2 Volume-Temperature Data for a Gas sample (at constant mass and pressure).

Volume (mL)	Temperature in Kelvin	V/T
600	300	2
500	250	2
400	200	2
300	150	2
200	100	2
100	50	2

Experiment 5.2



Effect of Temperature on the Volume of Gas

Objective: To observe the changes in volume of a gas as temperature changes.

Apparatus: Round bottomed flask, beaker, delivery tube and burner.

Procedure:

1. Set up the apparatus as shown in Figure 5.6.
2. Warm the flask gently with a low Bunsen flame.
3. Cool the flask and note what happens.
4. Record your observation.

Observations and analysis:

1. What do you observe from the experiment?
2. What is your conclusion from this activity?

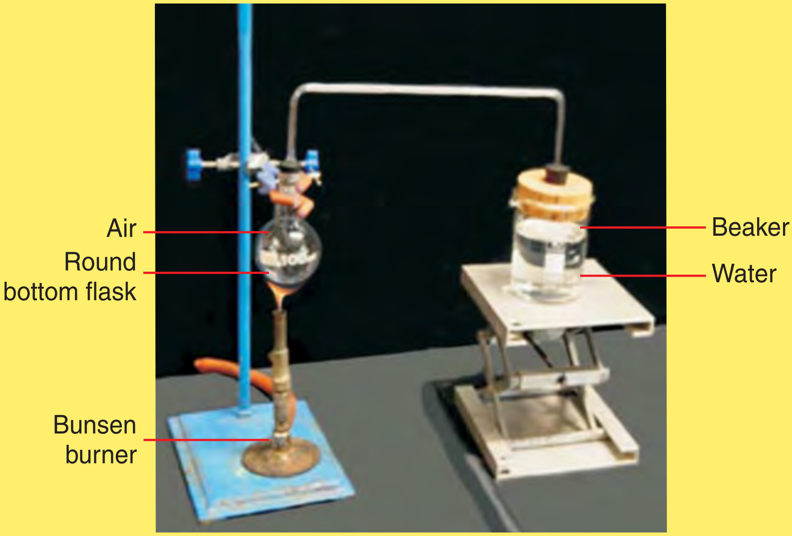


Figure 5.6 Relationship between temperature and volume of a gas.

Activity 5.10

Form a group and perform the following activity and present to the class.

Given the following data at a constant pressure:

Volume of Nitrogen gas (L)	Temperature (K)
4.28 L	303
5.79 L	410
7.77 L	550

a Draw a graph of the relationship between volume and temperature.

b Calculate the expected volume of the gas when the temperature reaches 700K.

c Explain the relationship between temperature and volume.

Example 2

A quantity of gas occupies a volume of 804 cm^3 at a temperature of 127°C. At what temperature will the volume of the gas be 603 cm^3, assuming that there is no change in the pressure?

Solution:

Given:

Initial conditions	Final conditions
$T_1 = 127^oC$ $= 127 + 273 = 400$ K $V_1 = 804\ cm^3$	$T_2 = ?$ $V_2 = 603\ cm^3$

Based on Charles' equation $\left(\frac{V_1}{T_1} = \frac{V_2}{T_2}\right)$, the unknown variable ($T_2$) can be calculated as:

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

$$= 400\,K \times \frac{603\,\cancel{cm^3}}{804\,\cancel{cm^3}} = 300\,K$$

Example 3

A gas at 65°C occupies 4.22 L. What will be the volume of the gas at a temperature of 36.9°C, assuming a constant pressure?

Solution:

Given:

Initial conditions	Final conditions
$T_1 = 65°C$ $= 65 + 273 = 338$ K $V_1 = 4.22$ L	$T_2 = 36.9°C + 273 = 309.9$ K $V_2 = ?$

From **Charles'** law

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

$$= \frac{4.22\ \text{L} \times 309.9\ \text{K}}{338\ \text{K}} = 3.87\ \text{L}$$

Exercise 5.5

1. At constant pressure, by what fraction of its volume will a quantity of gas change if the temperature changes from -173°C to 27°C?
2. At what temperature will the volume of a gas be
   a halved, b doubled,
   c tripled at constant pressure if the original temperature is 17°C?
3. At 25°C and 1 atm a gas occupies a volume of 1.5 dm^3. What volume will it occupy at 100°C and 1 atm?

3. *The Combined Gas Law*

A sample of a gas often undergoes changes in temperature, pressure, and volume. When this happens, the three variables must be dealt with at the same time.

Boyle's law and Charles' law can be combined to give one expression called the combined gas law. The combined gas law expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.

Derivation of the combined gas law:

Boyle's law: $V \propto 1/P$

Charles' law: $V \propto T$

Then, $V \propto T/P$ (*combined*)

$V = kT/P$ (where k *is a constant*)

It follows,

$$\frac{P_1V_1}{T_1} = k \quad \text{and} \quad \frac{P_2V_2}{T_2} = k$$

Since in each case k is constant, the combined gas law equation is given as follows:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Where P_1, V_1 and T_1 are the initial pressure, volume and temperature; P_2, V_2 and T_2 are the final pressure, volume and temperature of the gas respectively.

Example 4

A 300 cm^3 sample of a gas exerts a pressure of 60.0 kPa at 27°C. What pressure would it exert in a 200 cm^3 container at 20°C?

Solution:

Given:

Initial Conditions	V_1 = 300 cm^3	T_1 = 27 + 273 = 300 K	P_1= 60.0 kPa
Final Conditions	V_2 =200 cm^3	T_2 = 20 + 273 = 293 K	P_2 = ?

Using the combined gas law, $\left(\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}\right)$.

$$\Rightarrow \quad P_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{V_2} = \frac{60.0\,\text{kPa} \times 300\,\cancel{\text{cm}^3}}{300\,\cancel{\text{K}}} \times \frac{293\,\cancel{\text{K}}}{200\,\cancel{\text{cm}^3}}$$

P_2 = 87.9 kPa

Exercise 5.6

If a 50 cm^3 sample of gas exerts a pressure of 60.0 kPa at 35°C, what volume will it occupy at STP?

4. *Avogadro's law*

Activity 5.11

Form a group and discuss the following phenomena. Present your discussion to the class.

Suppose while you are playing a football in your school football team, the ball is accidentally deflated. Then immediately you fill the ball with air using air pump.

i. Why did the ball become strong enough?

ii. What happened to the number of particles in the ball?

iii. Which gas law can be obeyed? Explain.

The relationship between the volume of a gas and its number of molecules was explained by Avogadro. Avogadro's law states that *equal volumes of different gases, under the same conditions of temperature and pressure, contain the same number of molecules*. Thus, according to the law the volume of a gas is proportional to the number of molecules (*moles*) of the gas at STP.

Mathematically, $V \propto n$; where V is the volume and n is number of moles.

5. *The Ideal Gas Equation*

Activity 5.12

Form a group and discuss the following phenomenon. Present your discussion to the class.

A balloon can burst when too much air is added into it. Describe what happens to the pressure, volume, temperature, number of moles, and the balloon itself as it is inflated and finally bursts. Can you derive an equation which describes these relationships?

An ideal gas is a hypothetical gas that obeys the gas laws. Real gases only obey the ideal gas laws closely at high temperature and low pressure. Under these conditions, their particles are very far apart. The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.

We have considered the three laws that describe the behavior of gases as revealed by experimental observations:

Boyle's law: $V \propto \frac{1}{P}$ (*at constant* T *and* n)

Charles' law: $V \propto T$ (*at constant* P *and* n)

Avogadro's law: $V \propto n$ (*at constant* P *and* T)

This relationship indicates how the volume of gas depends on pressure, temperature and number of moles.

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

or $$V = R\frac{nT}{P}$$

where R, is a proportionality constant called the gas constant.

$$PV = nRT \quad (\textit{the ideal gas equation})$$

Thus, the ideal gas equation describes the relationship among the four variables P, V, T and n. An ideal gas is a gas whose pressure-volume-temperature behavior can be completely explained by the ideal gas equation.

At STP, the values of R can be calculated from the ideal gas equation.

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})\ (22.414\ \text{L})}{(1\ \text{mol})(273.15\ \text{K})}$$

$$= 0.082057 \frac{\text{L.atm}}{\text{K.mol}} = 8.314 \text{ L. kPa/K.mol} = 8.314 \text{ J/mol.K}$$

For calculations, we round off the value of R to three significant figures (0.0821 L.atm/K.mol) and use 22.4 L for the molar volume of a gas at STP.

Example 5

Calculate the volume (in liters) occupied by 7.4 g of CO_2 at STP?

Solution:

The ideal gas equation is given as

$$PV = nRT$$

$$V = \frac{nRT}{P} \left(\text{since } n = m/M \text{ by rearranging } V = \frac{mRT}{M\,P} \right)$$

$$= \frac{7.4 \cancel{\text{g}}}{44 \cancel{\text{g}}/\cancel{\text{mol}}} \times 0.082 \frac{\text{L.}\cancel{\text{atm}}}{\cancel{\text{K}}\ \cancel{\text{mol}}} \frac{273 \cancel{\text{K}}}{1 \cancel{\text{atm}}} = 3.77 \text{ L}$$

Example 6

At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of the gas.

Solution:

Given: V = 0.280 L, m = 0.400 g

At conditions of standard temperature and pressure

$$T = 273 \text{ K, P} = 1 \text{ atm}$$

$$R = 0.082 \text{ L.atm/K.mol}$$

$$\Rightarrow \quad PV = nRT$$

$$PV = \frac{m}{M} \cdot RT \ (\text{Since } n = \frac{m}{M});$$

$$\Rightarrow \quad M = \frac{mRT}{PV} = \frac{0.400 \text{ g} \times 0.082 \cancel{\text{L}}.\cancel{\text{atm}}/\cancel{\text{K}}.\text{mol} \times 273 \cancel{\text{K}}}{1 \cancel{\text{atm}} \times 0.280 \cancel{\text{L}}} = 31.98 \text{g/mol}$$

Exercise 5.7

The density of a gas at a pressure of 1.34 atm and a temperature of 303 K is found to be 1.77 g/L. What is the molar mass of this gas?

6. *Graham's Law of Diffusion*

Activity 5.13

Form a group and discuss the following phenomenon. Present your discussion to the class.

Explain why a helium-filled balloon deflated over time faster than an air-filled balloon does. (Hint: A balloon has many invisible pinholes).

We have seen that a gas tends to expand and occupy any space available to it. This spreading of gas molecules is called diffusion.

How do you compare the rate of diffusion of molecules with different densities?

Thomas Graham (1805 - 1869), an English chemist, studied the rate of diffusion of different gases. He found that gases having low densities diffuse faster than gases with high densities. He was able to describe quantitatively the relationship between the density of a gas and its rate of diffusion. In 1829, he announced what is known as Graham's law of diffusion.

Graham's law of diffusion states that *at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.*

Mathematically it can be expressed as:

$$r \propto \sqrt{\frac{1}{d}} \quad \text{or} \quad r \propto \sqrt{\frac{1}{M}};$$

where r is the rate of diffusion, d is the density and M is the molecular mass of the gas.

For two gases (Gas 1 *and* Gas 2), their rates of diffusion can be given as:

$$r_1 \propto \sqrt{\frac{1}{d_1}} \quad \text{or} \quad r_1 \propto \sqrt{\frac{1}{M_1}}$$

and

$$r_2 \propto \sqrt{\frac{1}{d_2}} \quad \text{or} \quad r_2 \propto \sqrt{\frac{1}{M_2}}$$

Rearranging these relationships gives the following expression

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad \text{or} \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}};$$

where r_1, d_1 and M_1 represent the rate of diffusion, density and molecular mass of gas 1. r_2, d_2 and M_2 represent the rate of diffusion, density and molecular mass of gas 2.

Experiment 5.3

Determination of Diffusion of Gases

Objective: To compare the rates of diffusion of HCl and NH_3 gases.
Apparatus: Glass tube, cork, cotton and stand.
Chemicals: conc NH_3, conc. HCl.

***Procedure*:**

1. Set up the apparatus as shown in Figure 5.7.
2. Insert pieces of cotton at the two ends of the tube.
3. Add 8 drops of concentrated ammonia on the cotton at one end and 8 drops of concentrated hydrochloric acid at the other end of the tube at the same time. Immediately close the two ends with corks.
4. Watch till a white ring is formed and record the time at which the white ring is formed.
5. Measure the distances between the white ring and the two ends.

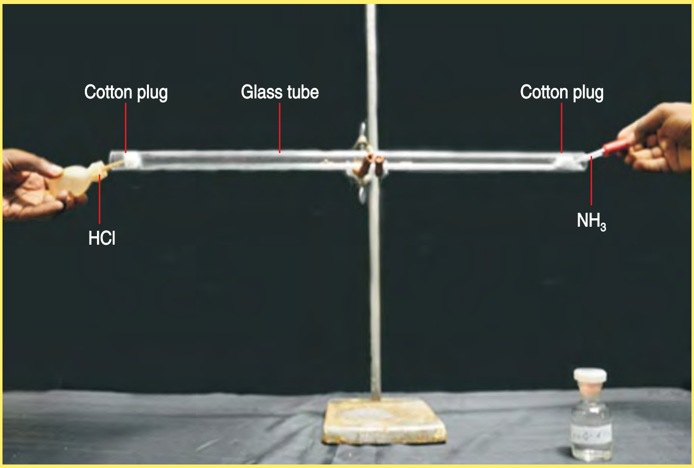


Figure 5.7 Determination of diffusion of gases.

Observations and analysis:

1. Which gas has moved the shorter distance to the white ring?
2. How do you compare the rate of diffusion of the two gases?
3. Which gas diffuses faster? HCl or NH_3?
4. Write your conclusions about the experiment.

Example 7

Which gas will diffuse faster, ammonia or carbon dioxide? What is the relative rate of diffusion?

Solution :

The molecular weight of CO_2 is 44 g/mol and that of NH_3 is 17 g/mol. Therefore, NH_3 diffuses faster than CO_2.

We can calculate the rate of diffusion as follows:

Let the rate of diffusion of NH_3 be r_{NH_3}

Let the rate of diffusion of CO_2 be r_{CO_2}

$$\frac{r_{NH_3}}{r_{CO_2}} = \sqrt{\frac{M_{CO_2}}{M_{NH_3}}} = \sqrt{\frac{44}{17}} = 1.6$$

This means rate of diffusion of NH_3 is 1.6 times that of CO_2.

Example 8

The rate of diffusion of methane (CH_4) is twice that of an unknown gas. What is the molecular mass of the gas?

Solution :

Let r_{CH_4} and r_x be the rates of diffusion of CH_4 and the unknown gas respectively.

Let M_{CH_4} and M_x be the molecular masses of CH_4 and the unknown gas respectively.

The rate of diffusion of CH_4 is two times faster than the unknown gas. This can be written mathematically as $r_{CH_4} = 2r_x$.

Now, substitute $2r_x$ in place of r_{CH_4} and solve for M_x using Graham's law.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \Rightarrow \frac{r_{CH_4}}{r_x} = \sqrt{\frac{M_x}{M_{CH_4}}}$$

$$\frac{2r_X}{r_X} = \sqrt{\frac{M_X}{16}} \Rightarrow M_X = 64$$

Therefore, the molecular mass of the unknown gas is 64.

Also the rate at which a gas diffuses is inversely proportional to the time taken. Mathematically,

$$r \propto \frac{1}{t}$$

If two different gases (gas 1 and gas 2) under the same conditions of temperature and pressure diffuse through a porous container, then the time required to diffuse for the two gases can be given by the following formula:

$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

where t_1 and t_2 are the time taken, r_1 and r_2 are the rates, M_1 and M_2 are the molecular masses of Gas 1 and Gas 2 respectively.

Activity 5.14

Form a group and discuss the kinetic molecular theory to explain the compression and expansion of gases. Present your findings to the class.

Reading Check

What names are given for the following ideal gas relationships?

a Volume and moles at constant temperature and pressure.

b Volume and pressure at constant temperature and moles.

c Volume and Kelvin temperature at constant pressure and moles.

5.4 THE LIQUID STATE

Competencies

By the end of this section, you will be able to:

- explain the terms: evaporation, boiling, condensation, vapor pressure; boiling point, molar heat of vaporization and molar heat of condensation;
- carry out an activity to demonstrate the concept of vapor pressure; and
- carry out an activity to determine the boiling points of water and ethanol.

Activity 5.15

Discuss the following activities in your group and present to the class.

1. Why some liquids are volatile and others are not?
2. What is the relationship between altitude and boiling point of a liquid?

You recall that liquids have a definite volume and an indefinite shape. They take the shape of their containers to the level they fill. On the average, liquids are more dense than gases, but less dense than solids.

As in a gas, particles in a liquid are in constant motion. However, the particles in a liquid are closer together than those in a gas. The attractive forces between particles in a liquid are more effective than between particles in a gas. This attraction between liquid particles is caused by the intermolecular forces (*dipole-dipole forces*, *London dispersion forces, and hydrogen bonding*).

Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of liquid particles. Accordingly, liquid particles are not bound together in fixed positions.

Energy Changes in Liquids

Activity 5.16

Form a group and discuss the following phenomenon:

When you take bath with hot water in your bathroom, the water collects on the mirror of the bathroom.

Present your discussion to the class.

The process by which a liquid changes to a gas is known as vaporization. Evaporation is the process by which liquid molecules break freely from the liquid surface and enter the vapor phase. Evaporation is explained in terms of the energy possessed by the molecules on the surface of the liquid.

In an open container, evaporation continues until all of the liquid enters the vapor phase. However, liquids in a closed container behave differently. The volume of the liquid decreases for a period of time, and remains unchanged. In closed containers, the vapor cannot escape. As the vapor concentration increases, some of the vapor molecules lose energy and return to the liquid state. When a vapor returns to the liquid state, it is said to condense. The process is called condensation. Evaporation and condensation are opposing processes.

$$\text{Liquid} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftharpoons}} \text{Gas}$$

The rate of evaporation of a liquid depends on three factors. These are temperature, intermolecular forces, and surface area of the liquid.

An increase in temperature increases the average kinetic energy of the molecules and this increases the tendency to change into the gaseous state. Some liquids evaporate readily at room temperature. Such liquids are said to be volatile. The volatile liquids have relatively weak forces of attraction between particles. Diethyl ether, ethyl alcohol, benzene and acetone are volatile liquids.

Non-volatile liquids have a little tendency to evaporate at a given temperature. They have relatively stronger attractive forces between their molecules, e.g., sulphuric acid, water, and molten ionic compounds.

Vapour pressure: The partial pressure of the vapour above a liquid is called vapour pressure. The vapour pressure of a liquid depends up on the temperature. At a given temperature, vapour pressure is constant. As the temperature increases, the vapour pressure of a liquid also increases due to high rate of evaporation.

Vapour pressure depends also on the strength of the intermolecular forces between the particles of the liquid. The stronger the intermolecular forces, the lower the vapour pressure will be, because fewer particles will have enough kinetic energy to overcome the attractive force at a given temperature. For example, water and alcohol have relatively low vapour pressure. On the contrary, liquids with low intermolecular forces have high vapour pressures at room temperature. For example, diethyl ether, a non-polar molecule with relatively weak dispersion forces, has a relatively higher vapour pressure.

Boiling and Boiling Point

Activity 5.17

Form a group and discuss the following ideas. Present your discussion to the class.

1. How does the kinetic-molecular theory explain why atmospheric pressure is greater at lower altitude than at a higher altitude?
2. Ice melts normally at 0°C. What happens to the melting point of ice in the presence of impurities? Does it melt below 0°C, above 0°C or exactly at 0°C? Explain.

What is the difference between evaporation and boiling? How does boiling point depend on the external pressure?

Boiling is the change of a liquid to bubbles of vapour that appear throughout the liquids. It is the conversion of liquid to vapour within the liquid as well as at its surface. It occurs when the equilibrium vapour pressure of the liquid equals the atmospheric pressure. During evaporation only molecules at the surface escape into the vapour phase, but at the boiling point the molecules within the liquid have sufficient energy to overcome the intermolecular attractive forces of their neighbors, so bubbles of vapour are released at the surface. It is the formation of vapour bubbles within the liquid itself that characterizes boiling and distinguishes it from evaporation.

If the temperature of the liquid is increased, the equilibrium vapour pressure also increases. Finally, the boiling point is reached. The boiling point of a liquid is the temperature at which the equilibrium vapour pressure of the liquid equals the atmospheric pressure. Therefore, the lower the atmospheric pressure, the lower the boiling point will be.

The boiling point of a liquid can be reduced as lowering the external pressure, because the vapour pressure of the liquid equals the external pressure at a lower temperature.

If the external pressure is 1.0 atmosphere (760 mmHg), the boiling point is called normal boiling point. For instance water boils, at 100°C at 1.0 atmospheric pressure. Thus, the normal boiling point of water is 100°C.

Where will water boil first? In Addis Ababa** or **Hawassa? Why?

Experiment 5.4

Observing the Vapour Pressure of liquid

Objective: To observe the vapour pressure of liquid.

Apparatus: Erlenmeyer flask, rubber bung, U-tube and burner.

***Procedure*:**

1. Set up the apparatus as shown in Figure 5.8.
2. Add about 100 mL of water into the Erlenmeyer flask and put a stopper. Heat the flask to expel the air above the water in the flask.
3. Half fill the U-tube with water.
4. Connect the U-tube to Erlenmeyer flask and note the water level in the two arms of the U-tube.
5. Heat the flask gently and observe the water level changes in the arms of the U-tube.

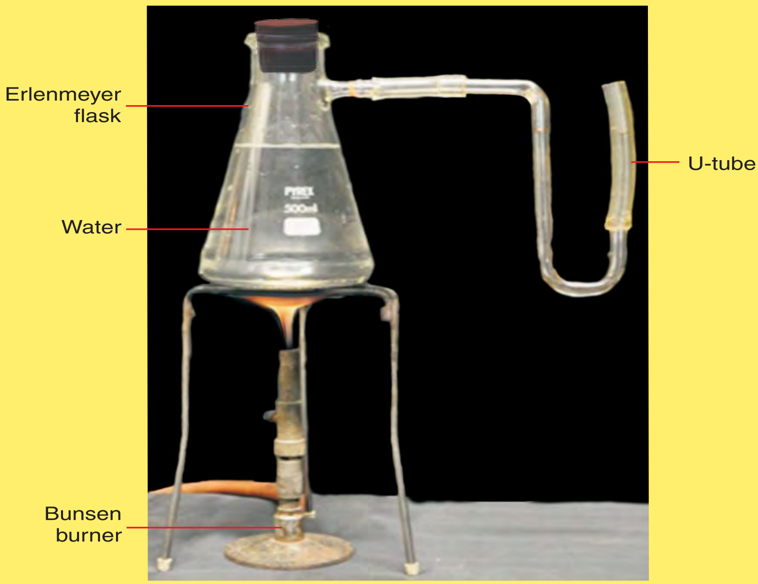


Fig. 5.8 Determination of vapour pressure.

Observations and analysis:

1. What do you observe? Give an explanation for the observation.

Experiment 5.5

Determining of Boiling Point

Objectives: To determine the boiling point of water.

Apparatus: Test tube, stopper, thermometer, beaker, burner, clamp, stand and base.

***Procedure*:**

1. Half fill the test tube with a sample of pure water and add some porclein chips.
2. Take a rubber stopper and pierce a thin opening on the side of the rubber stopper to allow the vapour to escape.
3. Fit the thermometer with the rubber stopper and insert it in the test tube.
4. Put the test tube in a beaker containing oil as shown in Figure 5.9.
5. Heat the oil in beaker gently and record the temperature at which the water boils.

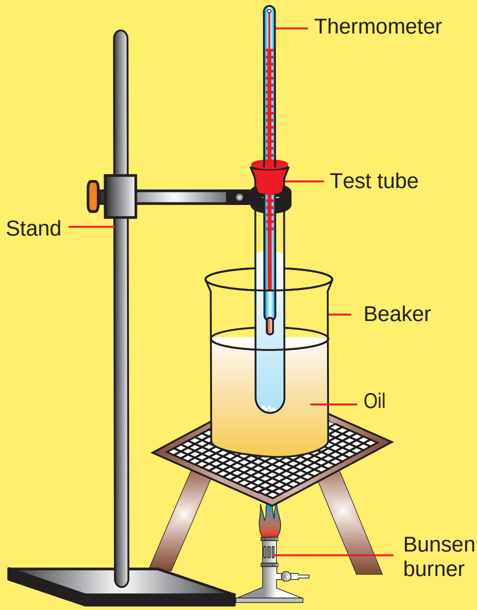


Figure 5.9 Determination of boiling point.

Observations and analysis:

1. What is the boiling point of water that you obtained from the experiment?
2. Does the temperature from the thermometer reading increase after the water starts to boil?
3. Explain why the thermometer was not put into the liquid.
4. Explain the purpose of adding porclein chips.

Boiling of a liquid requires a certain amount of heat energy to break the forces of attraction that holds the molecules together. The amount of heat energy necessary to bring about the vaporization of a fixed amount of a liquid at a fixed temperature to the gaseous state is called the heat of vaporization. For example, the heat of vaporization per mole of water at 298 K and 1 atmosphere is 44.0 kJ. This is called the molar heat of vaporization (ΔH_{vap}) of water.

Molar heat of vaporization (ΔH_{vap}) and molar heat of condensation (ΔH_{cond}) are equal in magnitude but opposite in sign, *i.e.*, $\Delta H_{vap} = -\Delta H_{cond}$. Vaporization is an endothermic processes where as condensation is an exothermic process.

Reading Check

1. What are the effects of impurity on the boiling point of liquids?
2. Why does the boiling point of liquid decrease as altitude increases?

5.5 THE SOLID STATE

Competencies

By the end of this section, you will be able to:

- explain the terms melting, fusion, sublimation, melting point, freezing point, molar heat of solidification;
- describe phase changes;
- explain temperature changes associated with phase changes;
- determine melting point of ice; and
- demonstrate an experiment to show the phase changes from ice to liquid water and then to water vapor.

Activity 5.18

Form a group and discuss the following:

1. When the crystals of iodine are warmed, they disappear into vapours without being changed into liquid.
2. When ethyl alcohol is taken in an open container it disappeares after sometime.

Present your discussion to the class.

When a solid is continuously heated its ordered crystalline structure is disturbed. The particles attain their freedom of motion gradually and melting (or fusion) takes place where the solid is converted into the liquid state. The temperature at which a crystalline solid is converted to a liquid is known as the melting point.

In contrast when a liquid is cooled, the molecules become closer and closer, and the intermolecular forces of attraction become stronger and stronger. The particles arrange themselves into a regular pattern in the liquid. This process is called freezing or solidification. Note that the freezing point of a liquid is the same as the melting point of a crystalline solid. This means that:

- Both melting and freezing take place at the same temperature.
- Both liquid and solid phases coexist at equilibrium with each other at the melting or freezing point.

To clarify this, let us consider ice, which melts at 0°C and water freezes at 0°C. Ice and water coexist in equilibrium with each other at 0°C as follows.

$$\text{Ice} \underset{\text{freezing}, 0°\text{C}}{\overset{\text{melting}, 0°\text{C}}{\rightleftharpoons}} \text{Liquid water}$$

The amount of heat needed to convert one gram of a solid to a liquid at the melting point is called heat of fusion. The molar heat of fusion or molar enthalpy of fusion (ΔH_{fus}) is the quantity of heat needed to convert one mole of a solid at its melting point to the liquid state. For example, the molar heat of fusion of ice is 6.01 kJ at 0°C. This is the amount of energy needed to break the attractive forces in the solid at the melting point. Melting requires the supply of energy; therefore, it is an endothermic process.

During the process of solidification, an amount of heat equal to the heat of fusion must be liberated. This quantity of heat liberated, which is exactly equal to the heat of fusion, is called the heat of solidification or heat of crystallization.

The molar heat of crystallization (ΔH_{cryst}) is the quantity of energy that must be removed from one mole of a liquid to convert it to the solid state at its freezing point.

$$\Delta H_{cryst} = -\Delta H_{fus}$$

Although the motion of the particles in a solid are more restricted than those in a liquid, many solids have a significant vapour pressure and evaporate directly from the solid to the vapour state without passing through the liquid state. This process is called sublimation.

$$\text{Solid} \underset{\text{deposition}}{\overset{\text{sublimation}}{\rightleftharpoons}} \text{Vapour (gas)}$$

During the sublimation process heat energy is absorbed. That is, it is an endothermic process.

Molar heat of sublimation (ΔH_{fus}) is the quantity of heat required to convert one mole of a solid to a gas at its sublimation point. The heat (enthalpy) of sublimation is related to the enthalpies of fusion and vaporization by:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

Phase Changes and Energy Changes in Solids

A phase is any part of a system that has uniform composition and properties. A state of matter represents a phase. Most solid substances undergo two changes of state when heated. A solid change to a liquid at the melting point, and the liquid changes to vapour at the boiling point. To understand state changes, we will consider the heating curve for a substance given in Figure 5.10.

A heating curve is a plot of temperature verses the uniform addition of heat. This can be illustrated for a hypothetical substance, in which the temperature of the substance is on the vertical axis and the passage of time during which heat is added to the substance is on the horizontal axis. Figure 5.10 shows the changes in the temperature and phases of a pure substance as it is heated, beginning with a solid and continuing to the gaseous state as described.

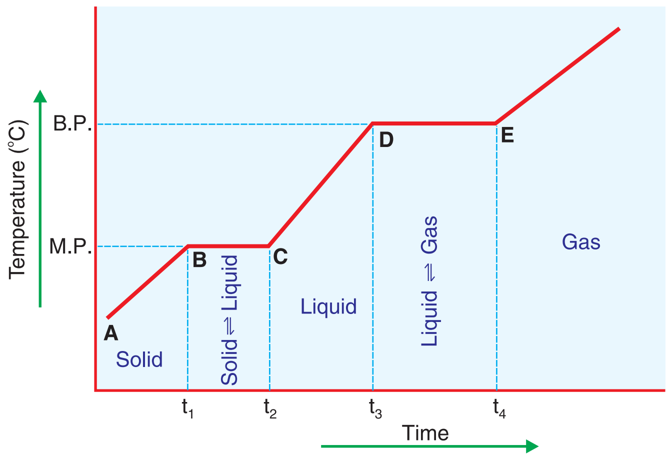


Figure 5.10 Heating curve.

Initially, the substance exists in the solid state, and the addition of heat increases its temperature. When the solid is heated, its temperature rises (A to B) until it reaches the melting point (point B), and the temperature remains constant (B to C) until all the solid is converted to a liquid (point C). The added heat energy is used to break the intermolecular forces, thus disrupting the solid structure. At point C phase change is completed. Once melting is completed, heating of the liquid raises its temperature (C to D) until the boiling point is reached at point D. In region (D to E) the addition of heat is utilized to break the intermolecular forces of the liquid to change it to a gas. When all the liquid has been converted into a gas, addition of heat simply raises the temperature of the gas.

Reading Check

Freezing and melting point of ice takes place at $0^{o}C$. Rationalize this.

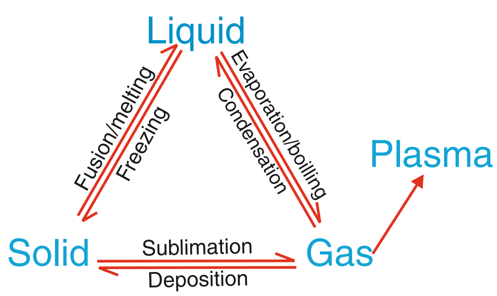
Check list

Key terms of the unit

- *Absolute zero*
- *Atmospheric pressure*
- *Avogadro's law*
- *Boiling*
- *Boiling point*
- *Boyle's law*
- *Charles' law*
- *Collision*
- *Combined gas law*
- *Condensation*
- *Evaporation*
- *Fluid*
- *Freezing*
- *Freezing point*
- *Gas constant*
- *Gas laws*
- *Gases*
- *Graham's law*
- *Heat of condensation*
- *Heat of vaporization*
- *Heating curve*
- *Ideal gas equation*
- *Kelvin temperature*
- *Kinetic molecular theory of gases*
- *Liquids*
- *Melting*
- *Melting point*
- *Molar heat of condensation*
- *Molar heat of fusion*
- *Molar heat of sublimation*
- *Non-volatile liquid*
- *Normal boiling point*
- *Phase change*
- *Properties of gases*
- *Solidification*
- *Standard temperature and pressure (STP)*
- *Sublimation*
- *Vapor pressure*
- *Volatile liquid*

Unit Summary

- *Matter is anything that has mass and occupies space.*
- *Matter exists in the form of a solid, a liquid or a gas.*
- *Solids have a definite volume and a definite shape.*
- *A liquid has no definite shape, it takes the shape of its container.*
- *A gas has neither a definite volume nor a definite shape.*
- *Gases and liquids are fluids where as solids are not. Fluidity is the tendency to flow.*
- *When energy is supplied to a solid, it melts and changes to a liquid, the particles move faster. Additional energy will make the liquid boil and form a gas.*
- *In the gas the particles are much more widely spaced and move much faster than in liquid and solid.*
- *Some solids directly change to gases by the process of sublimation.*
- *Phase change can be illustrated as:*



- *General (combined) gas equation:* $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- *Ideal gas equation:* $PV = \mathrm{n}RT$
- *Grahm's Law of diffusion:* $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$
- *Melting point = Freezing point*
- *Boiling point = Condensation point*
- *Sublimation point = Deposition point.*

REVIEW EXERCISE ON UNIT 5

Part I: Assert your answers based on the given instructions:

We have three physical states of substances namely solid, liquid, and gas. Identify the following as solid, liquid or gas based on the properties given to explain their characteristics.

1. They have a definite shape and a definite volume.
2. Their molecules are highly disordered.
3. The motion of their molecules is highly restricted.
4. They can be easily compressed.
5. They have a tendency to flow.
6. They can take the shape of their container.
7. They can move in all direction at high speed.
8. They can easily diffuse through each other.
9. They can sublime.
10. They have less density relative to the other states.

Part II: Matching-type questions

	A		B
11.	Melting point	A	Solid $\rightarrow$ gas
12.	Heat of fusion	B	Liquid $\rightarrow$ gas
13.	Heat of sublimation	C	Gas $\rightarrow$ solid
14.	Sublimation point	D	Solid $\rightarrow$ liquid
15.	Melting	E	Liquid $\rightarrow$ solid
16.	Freezing	F	Same as freezing point
17.	Sublimation	G	Same as deposition point
18.	Deposition	H	Same as heat of crystallization
		I	Similar to fusion
		J	Similar to crystallization
		K	Similar to heat of deposition

Part III: Write the missing words in your exercise book

19. In a solid the particles are very close together and can only _______ about a fixed position.
20. Sublimation occurs when a solid changes directly to a _______ passing the _______ state.
21. The melting point of a solid is the same as _______.
22. When the water boils, its vapour pressure is equals to _______.
23. The temperature at which a crystalline solid is converted to a liquid, is called ______.
24. The lowest attainable temperature is ______.
25. "Equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules"; this is a statement of _______ law.

Part IV: Short-answers type questions

26. What is the difference between volatile and non-volatile substances? Give an example of each.
27. Why is the temperature of a substance constant at its melting point; even though heat is added to it?
28. What factors affect the rate of evaporation of a liquid?

Part V: Problems to solve

29. Convert the following pressure measurements:
    - a 720 mmHg to atm
    - b 1.25 atm to mmHg
    - c 542 mmHg to atm
    - d 740 mmHg to kPa
    - e 700 kPa to atm
30. A 2.50 L container is filled with 175 g of argon:
    - a if the pressure is 10.0 atm, what is the temperature?
    - b if the temperature is 22 K, what is the pressure?
31. If 0.500 mole of nitrogen gas occupies a volume of 11.2 L at $0^{\circ}C$; what volume will 2.00 mole of nitrogen gas occupy at the same temperature and pressure?
32. A certain gas is found in the exhaust of automobiles and power plants. Consider a 1.53 L sample of a gas at a pressure of 5.6×10^4 Pa. If the pressure is changed to 1.5×10^4 Pa at a constant temperature, what will be the new volume of the gas?

Part III: Write the missing words in your exercise book

19. In a solid the particles are very close together and can only _______ about a fixed position.
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22. When the water boils, its vapour pressure is equals to _______.
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