Chapter # 04

States of Matter I: Gases

Major Concepts

4.1 Kinetic Molecular Theory of Gases
4.2 Boyle's Law
4.3 Charles's Law
4.4 Avogadro's Law
4.5 Ideal Gas Equation
4.6 Deviation from Ideal Gas Behavior
4.7 Van der Waal's Equation
4.8 Dalton's Law of Partial Pressure
4.9 Graham's Law of Diffusion and Effusion
4.10 Liquefaction of Gases
4.11 Fourth State of Matter: Plasma

Learning Outcomes

The students will be able to:

- List the postulates of Kinetic Molecular Theory. (Remembering)
- Describe the motion of particles of a gas according to Kinetic Theory. (Applying
- State the values of standard temperature and pressure (STP). (Remembering)
- Relate temperature to the average kinetic energy of the particles in a subst (Applying)
- Use Kinetic Theory to explain gas pressure. (Applying)
- Describe the effect of change in pressure on the volume of gas. (Applying)
- Describe the effect of change in temperature on the volume of gas. (Applying)
- Explain the significance of absolute zero, giving its value in degree Celsi Kelvin. (Understanding)
- State and explain the significance of Avogadro's Law. (Understanding)
- Derive ideal gas equation using Boyle's, Charles', and Avogadro's L (Understanding)
- Explain the significance and different units of ideal gas constant. (Understandi
- Distinguish between ideal and real gases. (Understanding)
- Explain why real gases deviate from the gas laws. (Analyzing)
- Define and describe the properties of Plasma. (Applying)
- Derive new form of Gas Equation with volume and pressure corrections fo gases. (Understanding)

- State and use Graham's Law of diffusion. (Understanding)
- State and use Dalton's Law of Partial Pressures. (Understanding)
- Describe some of the implications of the Kinetic Molecular Theory, such as the velocity of molecules and Graham's Law. (Applying)
- Explain Lind's method for the liquefaction of gases. (Understanding)
- Define pressure and give its different units. (Remembering)
- Define and explain plasma formation. (Understanding)

Introduction

What is matter? Matter is anything which has mass and occupies space. Matter is classified by its physical state as solids, liquids, gases and plasma. Under normal conditions, most matter on earth exists in one of three physical states, namely, solid, liquid, or gas. Solids have both definite volume and definite shape. Their particles are close together and fixed into place due to the greatest interaction forces. Liquids have definite volume but have no definite shape. They adopt the shape of container in which they are placed. Their particles are still close together but move freely. Their molecules have greater attractive forces than gases but less than solids. Gases neither have definite volume nor definite shape. They occupy all the available space of the container. Their molecules are neither close together nor fixed in place due to weak attractive forces. The gases occupy about 0.1% (99.9% empty space) whereas liquids and solids occupy about 70% (30% empty space) of the total volume at room temperature and one atmosphere pressure. Our body contains matter in all three physical states: solid (bone, hair, muscles etc.), liquid (blood and body fluids), and gas (the gases we inhale and exhale and the gases that are produced in the intestine). The particular state of a substance is determined by temperature and pressure under which they exist. For example water changes to solid (ice or snow) at/below 0°C and changes to gas (steam) above 100°C. The physical state of a substance usually means its state at room temperature (about 20°C to 25°C) and at

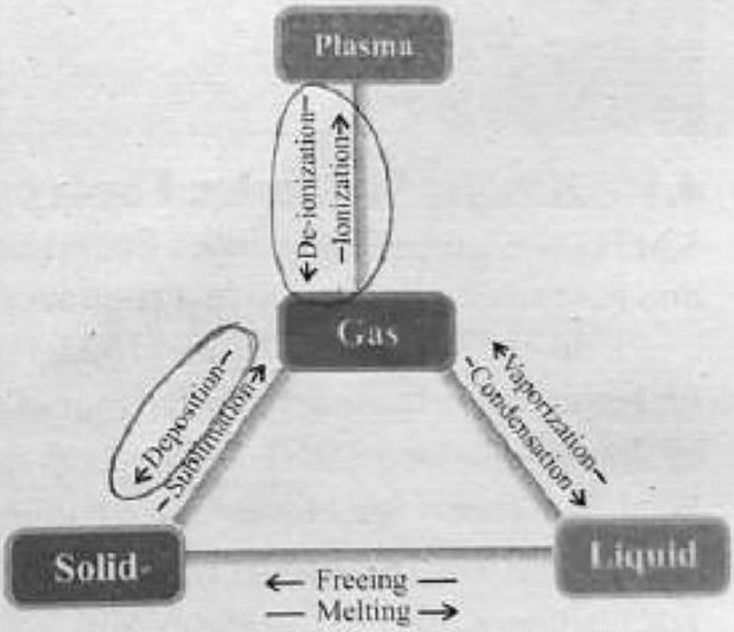

Figure 4.1: Diagram showing possible changes of states of matter

one atmosphere pressure. At room temperature and one atmosphere pressure, physical state of water, milk and vinegar is liquid; salt (NaCl), sugar (sucrose), gold is solid; and that of oxygen, carbon dioxide and methane is gas. The simple form of matter is the gaseous state. The most of the matter around us is in the s state. The liquid state of any substance exists only with in a relatively narrow ra of temperature and pressure. A fourth state of matter, called plasma, is less famil Plasma is ionized gas mixture. Stars are made of plasma. Fire is an example plasma. Most of the universe is composed of plasma. Plasma, like gases have indefinite shape and an indefinite volume.

Out of 118 elements only 11 elements are gases under normal conditions a they are hydrogen, helium, nitrogen, oxygen, fluorine, neon, chlorine, argo krypton, xenon and radon.

Table 4.1: Some Compounds Found as Gases at Room Temperature (25°C) and One Atmosph Pressure

HCl	Hydrogen chloride	CO	Carbon monoxide
HBr	Hydrogen bromide	CO_2	Carbon dioxide
NO	Nitric oxide	NH_3	Ammonia
N_2O	Nitrous oxide	NO_2	Nitrogen dioxide
H_2S	Hydrogen sulphide	SO_2	Sulphur dioxide
HCN	Hydrogen cyanide	SO_3	Sulphur trioxide

4.1 Kinetic Molecular Theory of Gases (KMTG)

KMTG was proposed by Swiss Mathematician Bernoulli. In 1738, he suggest that gaseous molecules are in a continuous state of motion.

In 1857, Rudolf Clausius (1822-1888) postulated 'Kinetic molecular the for better understanding the behaviour of gases. This theory was further develo by James Maxwell (1831-1879), Ludwig Boltzmann (1844-1906), and Van Waal. This theory sometimes called Billiard ball model.

4.1.1 Postulates of Kinetic Molecular Theory

The main postulates of this theory are:

i) All gases consist of very large number of tiny particles called atoms molecules.

ii) The molecules of gases are widely separated from each other. Hence mos the volume of the gas is empty space (nearly 99.96%). That is why gases ca compressed and mixes completely with each other.

iii) The actual volume of gas molecules is negligible (very small) as compar the volume of the gas or volume of its container.

iv) Gas molecules are in constant random motion in straight lines. Their direction of motion changes only when they collide with one another or with the walls of container. That is why gases quickly and completely fill any container in which they are placed. Pressure exerted by gas is due to collision of molecules with one another or with the walls of container.

v) The collisions of gas molecules are completely elastic. This means that total energy of molecules before and after the collision remains same. In other words, energy can be transferred from one molecule to another as a result of a collision. However, the total energy of all the molecules in a system remains the same. Therefore, the average kinetic energy of gas molecules is not affected by these collisions and remains constant as long as there is no change in temperature.

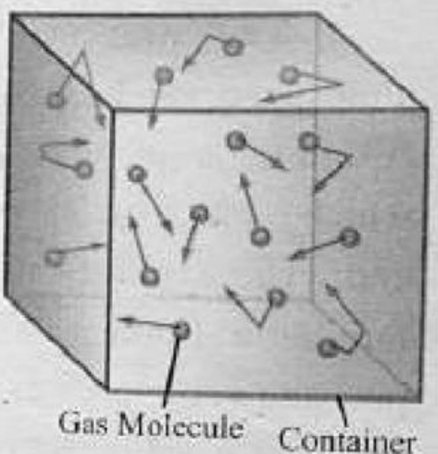

Figure 4.2: Random motion of gas molecules

vi) Due to large intermolecular distance, no attractive or repulsive forces are present among gas molecules. Therefore, they are independent in their behaviour.

vii) Gas molecules have kinetic energy. This is due to constant motion. By increasing temperature, the average kinetic energy of gas molecules increases. Any two gases at the same temperature will have the same average kinetic energy. Gas particles move faster as the temperature increases, hitting the walls of the container with more force and producing higher pressures.

Interesting Information:

In hot summer days, oxygen gas molecules move faster than a jet airplane at an average speed of about 1660 Km per hour. But those oxygen gas molecules are unable to cross the room very fast because each molecule of gas has about 5 to 6 billion collisions with other molecules in one second.

4.1.2 Pressure and its Units

One of the most important properties of any gas is its pressure. Gas pressure is a result of the force exerted by the collision of particles with the walls of the container.

The force exerted by gas molecules per unit area per second is called gas pressure or simply pressure.

$$\text{Pressure(P)} = \frac{\text{Force (F)}}{\text{Area (A)}}$$

The pressure exerted by Earth's atmosphere can be measured with a barometer. A simple barometer is one meter long glass tube which is closed at one end and open at the other end. The barometer is probably the most familiar instrument for measuring atmospheric pressure (The force exerted by atmosphere on unit area is called atmospheric pressure). How atmospheric pressure is measured?

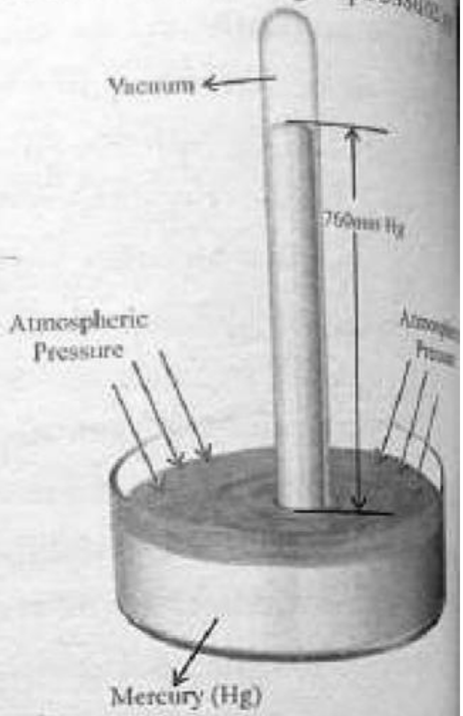

Figure 4.3: The height of mercury column in the barometer

Barometer is constructed by filling a glass tube with liquid mercury (Hg), and then inverting it into a mercury containing dish. The mercury in the tube on inverting, flow out in to the container, creating a vacuum the top (closed end) of the tube due to the force of gravity. This fall becomes equal the force exerted by the atmosphere which balances the gravitational force is cal atmospheric pressure.

The length of tube (column) measures atmospheric pressure. Atmospheric pressure varies from place to place and with changing weather condition. Therefore some standard pressure must be taken. The standard atmospheric (1 atm) pressure is the force exerted by 760 mm of mercury on an area of $1 cm^2$ at 0°C, the average atmospheric pressure at sea level.

Units

A number of different units are used to measure pressure. The common unit pressure is millimeters of mercury, symbolized mmHg. The pressure millimeter of mercury (mmHg) is also called the torr in the honor of Evangelista Torricelli (1608-1647), the Italian mathematician and physicist who invented barometer in 1643. Thus we have the following relation:

1 atm = 760 mmHg = 760 torr

The SI unit of pressure is the *Pascal (Pa)*, named in honor of mathematician and physicist Blaise Pascal (1623-1662). He was the first to propose that atmospheric pressure decreases with increasing altitude. One Pascal (Pa) is defined as the pressure exerted by a force of one Newton (1N) acting on an area of one square meter (m^2).

The commonly used units of pressure other than Pascal are as follows:

(i) mmHg	(ii) cm Hg	(iii) torr
(iv) atmosphere	(v) psi	(vi) bar

- 1 atm = 760 mmHg = 760torr = 76 cmHg
- 1 atm = 101325 P_a or Nm^{-2} = 101.325 kP_a
- 1 atm = 1.01325bar = 1013.25 mbar (millibar)
- 1 atm = 14.7psi (pounds per square inch)
- 1 atm = 29.92 inches of Hg

Torr is often used for low pressure. In the English or British system, pressure is measured in terms of pounds per square inch.

Keep in Mind

Barometer can be used only to measure atmospheric pressure and vapour pressure of a liquids. It cannot measure air pressure inside a tire and in a closed container like cylinder of gas or football. The pressure of the air inside an inflated tire or a football is measured by a tire-pressure gauge. The blood pressure is measured by sphygmomanometer.

Example 4.1

The air pressure inside the tire of a small car is 35psi. What is the equivalent pressure in (a) atm, (b) Pascal, (c) kPa, (d) bar and (e) torr?

Solution:

a) We know that, 1atm = 14.7psi

Hence, 14.7psi = 1atm

$$1\text{psi} = \frac{1\text{atm}}{14.7\text{psi}}$$

$$35\text{psi} = \frac{1\text{atm}}{14.7\text{psi}} \times 35\text{psi}$$

$$= 2.38\text{atm}$$

b) We know that, 14.7 psi = 1 atm = 101325Pa

Hence, 14.7 psi = 101325 Pa

$$1 \text{ psi} = \frac{101325\text{Pa}}{14.7\text{psi}}$$

$$35 \text{ psi} = \frac{101325\text{Pa}}{14.7\text{psi}} \times 35\text{psi}$$

$$= 241250\text{Pa}$$

c) We know that, 14.7 psi = 1 atm = 101.325kPa

Hence, 14.7 psi = 101.325kPa

$$1 \text{ psi} = \frac{101.325\text{kPa}}{14.7\text{psi}}$$

$$35 \text{ psi} = \frac{101.325\text{kPa}}{14.7\text{psi}} \times 35\text{psi}$$

$$= 241.250\text{kPa}$$

d) We know that, 14.7 psi = 1 atm = 1.01325 bar

Hence, 14.7 psi = 1.01325 bar

$$1 \text{ psi} = \frac{1.01325\text{bar}}{14.7\text{psi}}$$

$$35 \text{ psi} = \frac{1.01325 \text{ bar}}{14.7\text{psi}} \times 35\text{psi}$$

$$= 2.41250 \text{ bar}$$

e) We know that, 14.7 psi = 1 atm = 760 torr

Hence, 14.7 psi = 760 torr

$$1 \text{ psi} = \frac{760\text{torr}}{14.7\text{psi}}$$

$$35 \text{ psi} = \frac{760 \text{ torr}}{14.7\text{psi}} \times 35\text{psi}$$

$$= 1809.52\text{torr}$$

Practice Exercise 1:

The pressure of atmosphere on the top of Mt. Everest is 270 torr. Represent the pressure in **(a)** atm, **(b)** mbar, and **(c)** psi.

4.2 Boyle's Law(Pressure-Volume Relationship)

In 1662, the pioneering English scientist Robert Boyle and his assistant Robert Hooke (1635 - 1703) used a J-shaped glass tube closed at one end to measure the volume of gas at different pressures. They kept temperature constant. They observed an inverse relationship between volume and pressure-an increase in one results in a decrease in the other.

Robert Boyle (1627-1691)

This law can be stated in two ways:

i) The volume of a fixed amount of gas is inversely proportional to pressure at constant temperature.

If the gas pressure is doubled, the volume is halved; if the pressure is halved, the gas volume doubles.

$V \propto \frac{1}{P}$ (at Constant Temperature)

$V = K\frac{1}{P}$ Where K is proportionality Constant

or

$PV = K$

With the help of this equation, the Boyle's law can be defined as:

ii) At constant temperature, the product of pressure and volume is constant for the given mass of gas.

Mathematically,

$$P_1V_1 = P_2V_2 = P_3V_3 = K$$

It means that, P_1 is the initial pressure and V_1 is the initial volume. If we increase pressure to P_2 the volume will reduce to V_2 and if we again increase pressure to P_3, the volume will reduce to V_3 and so on.

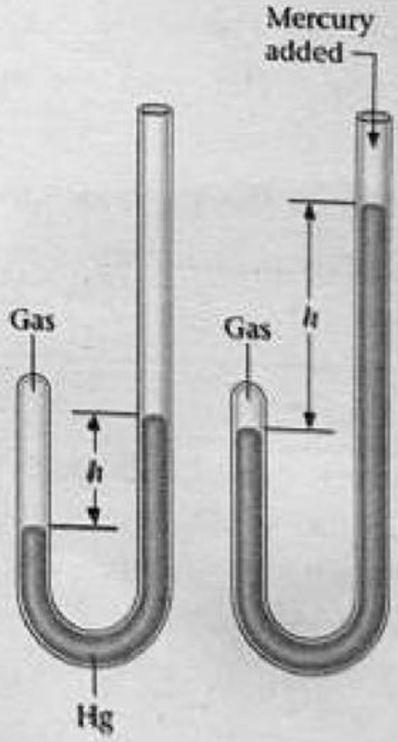

Figure 4.4: Verification of Boyle's Law

Conceptual Check Point:

What happens when you squeeze a hydrogen-filled balloon?

Point of Interest:

When a student of your class fills helium gas inside the balloon, it inflates, until the pressure of gas inside the balloon becomes equal to external pressure. Because the mass of same volume of helium is less than the mass of same volume of air (i.e. helium is less dense than air) at constant temperature and pressure. Because of this, it rises up. The balloon carry on travelling until the pressure inside and outside becomes equal. Then it stops rising and remains in one place in the atmosphere.

Graphical Representation of Boyle's Law

i) If a graph is plotted between pressure on the x-axis (abscissa) and volume on the y-axis (ordinate), then a curve is obtained, which shows that, pressure is inversely proportional to volume. This curve is called isotherm or hyperbola. (Iso means same; therm means heat).

ii) The process in which the temperature of the system remains constant is called isothermal process.

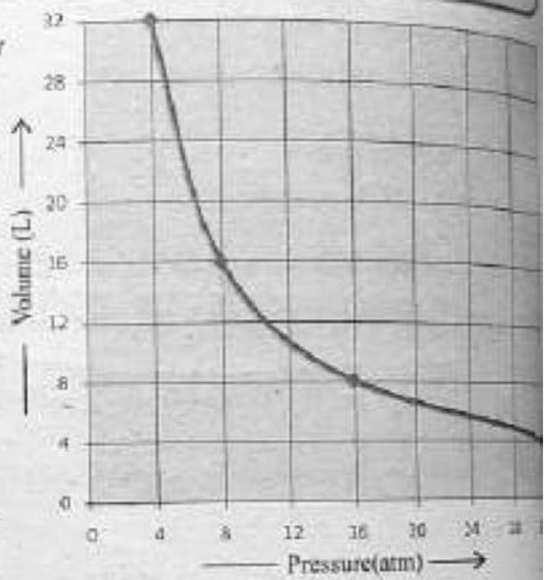

Figure 4.5:
Graphical representation of Boyle's la

Example 4.2

A green colour balloon occupies a volume of $25 dm^3$ when the gas pressure is 1.25atm. What will be its volume if the pressure is increased to 2.50atm at constant temperature?

Solution:

Initial volume $= V_1 = 25\ dm^3$

Initial pressure $= P_1 = 1.20$ atm

Final volume $= V_1 = ?$

Final pressure $= P_1 = 2.50$ atm

According to Boyle's law:

$$P_1V_1 = P_2V_2 \quad \text{or} \quad V_2 = \frac{P_1V_1}{P_2}$$

By putting the values we get,

$$V_2 = \frac{(1.25\ atm)(25\ dm^3)}{2.50\ atm}$$

$V_2 = 12.5\ dm^3$
By increasing pressure volume of gas decreases, hence final volume is less than initial volume.

Practice Exercise 2:

Methane gas found in Sui Gas (Natural gas) contains a volume of $70dm^3$ at a pressure of 75atm. What volume would the gas occupy if the pressure is increased by four times? The temperature of gas remained unchanged.

3 Charles's Law (Temperature-Volume Relationship)

Jacques Charles
(1746-1823)

1787, a French physicist Jacques Charles, who was the st person to fill a balloon with hydrogen gas and who ade the first solo balloon flight, found a relationship etween volume and temperature of a gas (volume increases temperature increases and decreases as temperature ecreases). This relation is called Charles law. In 1848, a ottish physicist William Thomson (1824-1907), whose e was Lord Kelvin, proposed an absolute-temperature ale, now known as the Kelvin scale. He identified 73.15°C as absolute zero, theoretically the lowest attainable temperature. In ms of the Kelvin scale, Charles's law states: Volume of a fixed amount of gas is ectly proportional to absolute temperature at constant pressure.
athematically,

$V \propto T$ (At constant pressure)

$V = K.T$

Where K is proportionality constant

$\frac{V}{T} = K$

If temperature is changed from T_1 T_2 and volume changes from V_1 to V_2, en:

$$\frac{V_1}{T_1} = K \quad \text{and} \quad \frac{V_2}{T_2} = K$$

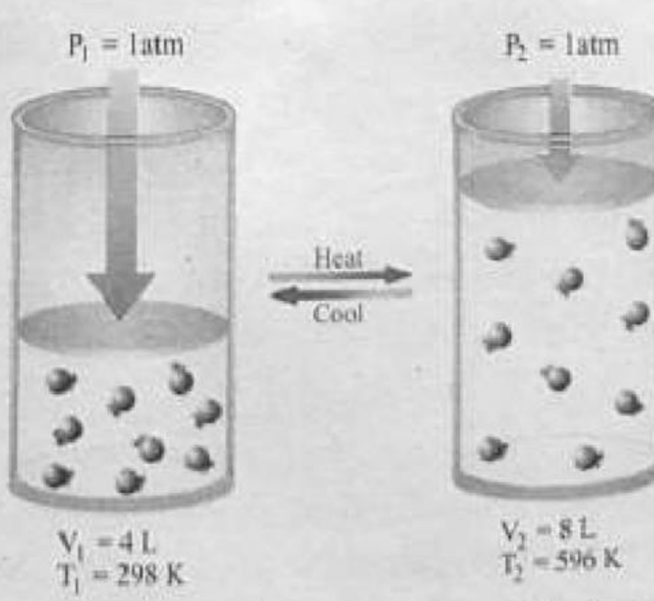

Figure 4.6: Verification of Charles's Law

So,

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

The law can also be stated as:

For the given mass of a gas, the ratio of volume and temperature remains constant at constant pressure.

Keep in Mind

The word *degree* and the symbol for degrees (°) are not used with the Kelvin scale. The Kelvin (absolute) scale must be used in all gas law problems involving temperature.

Conceptual Check Point:

- **What happens if someone put hot carbon dioxide filled-balloon into a freezer?**
- **We can put more gas during the cold winter days in the CNG cylinder of car as compared to hot summer days, why?**
- **Have you ever noticed that the same amount of gas exerts more pressure inside the LPG cylinder at Sibi and less pressure at Ziarat?**

Graphical Explanation of Absolute Zero:

When a graph is plotted between temperature on x-axis and volume on y-axis for the sample of a gas, a straight line is obtained which intersects the temperature axis at −273.15°C. This shows that if the sample of gas is cooled to −273.15°C, the volume of the gas will become zero. Any volume less than zero are not possible. The temperature at zero volume is −273.15°C (0K), no matter what kind of gas is used. This hypothetical temperature (−273.15°C) at which volume of a gas is considered as zero is called absolute zero or zero of Kelvin or absolute scale. This lowest possible temperature could be achieved only if the substance remains in gaseous state. Since all real gases are converted into liquid state before this temperature, therefore, this is not possible for real gases. This is against the Law of Conservation of Mass.

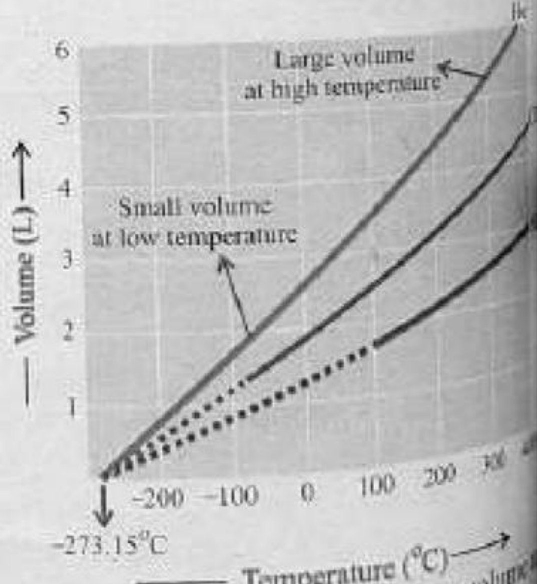

Figure 4.7: The graph between volume and temperature for three different gases

Keep in Mind

- Kelvin and a degree Celsius have the same magnitude. Thus, while we add 273.15 to the temperature in °C to get the temperature in Kelvin, a change in temperature in Celsius is equal to the change in temperature in Kelvin. A temperature of 25°C is the same as 298.15K. A change in temperature of 25°C, however, is the same as a change in temperature of 25 K.
- In most calculations we will use 273 instead of 273.15 as the term relating K and °C. By convention, we use **T** to denote absolute (Kelvin) temperature and **t** to denote Celsius temperature.

Absolute zero can also be defined as: the temperature at which the molecular motion ceases (stops) and a substance would have zero kinetic energy. At this temperature, none of the particles would be moving at all. Their speed and their kinetic energy would both be zero. The absolute zero is difficult to attain. Anyhow, recently the temperature as below as 10^{-6} K (0.000001K) has been attained which is greater than absolute zero. Charles law is not obeyed when temperature is in degree Celsius (°C). The temperature on Celsius scale has negative values (below 273.15 Kelvin). If we put negative values of temperature in to the equation, we get negative volumes. The negative volume for matter (gas) is not possible. In order to ensure that only values of V= 0 occur, we have to use an absolute temperature scale where T=0. The standard absolute scale is the Kelvin (K) scale.

Greater the mass (no. of moles) of a gas taken, greater would be the slope of the straight line. This is because, greater the number of mole of a gas, greater would be the volume of a gas.

Keep in Mind

The Scientists prefer the Kelvin scale to the Celsius, and certainly to the Fahrenheit scales. If the Kelvin temperature of a gas is doubled, its volume has doubled as well. The same cannot be said if the temperature were doubled from, say, 25°C to 50°C, or from 77°F to 122°F, since neither the Celsius nor the Fahrenheit scale can be used to calculate the volume of gas.

Example 4.3

A balloon is inflated to a volume of 1.5 dm^3 at room temperature (25°C). What would be the new volume of the balloon when it is placed in a hot water (temperature 60°C)? The pressure stays constant.

Solution:

Initial volume $= V_1 = 1.5dm^3$

Initial temperature $= T_1 = 25^{\circ}C + 273 = 298K$
Final volume $= V_2 = ?$
Final temperature $= T_2 = 60^{\circ}C + 273 = 333K$
Pressure $= P = \text{Constant}$
According to Charles' law,

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

By putting the values we have,

$$V_2 = \frac{1.5dm^3}{298K} \times 333K$$

$$V_2 = 1.68dm^3$$

The final volume is greater than initial volume, because, according to Charles' law, volume increases by increasing Kelvin temperature.

Practice Exercise 3:

A child blows a bubble that has a volume of 25 cm^3 at 10°C. As the bubble rises, it comes across the hot air where the temperature is 30°C. What is the volume of the bubble, if the pressure does not change?

4.4 Avogadro's Law (Amount-Volume Relationship)


Amedeo Avogadro (1776-1856)

The Italian scientist Amedeo Avogadro (1776-1856), a professor of higher physics at the University of Turin for many years, observed a relationship between the volume and number of moles of a gas at constant temperature and pressure. This relation is known as Avogadro's law. This law was proposed in 1811 but it was not generally accepted until after 1858, when an Italian chemist, *Stanislao Cannizzaro*, set up a logical system of chemistry based on it. The law is approximately valid for real gases at sufficiently low pressures and high temperatures.

Avogadro's law can be stated in three ways:

i) Equal volumes of different gases contain equal numbers of molecules at the same temperature and pressure.

One dm^3 of hydrogen gas contain the same number of moles (or molecules) as one dm^3 of helium, nitrogen, oxygen or any other gas at fixed temperature and pressure.

(ii) Equal numbers of molecules of different gases occupy equal volumes at the same temperature and pressure.

One mole (6.02×10^{23} molecules) of hydrogen gas occupies the same volume ($22.414 dm^3$) as a one mole of oxygen, argon or any other gas at STP.

Table 4.2: Molar Volumes of Some Gases at STP.

Gas	Formula	Molar Mass (g/mol)	Molar Volume (dm^3/mol)	Number of Molecules in one mole
Hydrogen	H_2	2.02	22.428	6.02×10^{23}
Helium	He	4.002	22.426	6.02×10^{23}
Nitrogen	N_2	28.01	22.404	6.02×10^{23}
Oxygen	O_2	32.00	22.394	6.02×10^{23}
Chlorine	Cl_2	70.91	22.063	6.02×10^{23}
Ammonia	NH_3	17.03	22.094	6.02×10^{23}
Carbon dioxide	CO_2	44.01	22.256	6.02×10^{23}
Any ideal gas	------	------	22.414	6.02×10^{23}

We may say that,

The volume occupied by one mole of an ideal gas at STP is called molar volume. It is equal to $22.414 dm^3$. It is always constant.

Equal volumes of all ideal gases have equal number of molecules but different masses.

Mass and size of gas molecules do not affect volume of the gas.

Avogadro's law can also be stated as:

The volume of a gas is directly proportional to number of moles at constant ...erature and pressure.

According to Avogadro's law, the volume of an ideal gas at a constant ...ure and temperature depends on its molar amount. If the amount of the gas is ...ed, the gas volume is doubled; when the amount of the gas is tripled, the ...e is tripled; if the amount is halved, the volume is halved; when the amount of ... is zero, the volume is zero.

...athematically,

$V \propto n$ (at constant temperature and pressure)

where n represents the number of moles and **K** is the proportionality constant.

$$V = Kn \quad \text{or}$$

$$\frac{V}{n} = K$$

Have you ever inflated a balloon? If yes, then you experience Avogadro's law. With each exhaled breath, you add more gas particles to the inside of the balloon, increasing its volume. The more air you would put into the balloon, the greater its volume would be, because the volume of the balloon is the volume of the trapped gas. If a hole is poked in the balloon, the gas escapes, decreasing both the amount of gas and the volume.

For two samples of gas at the same temperature and pressure, the relation between volumes and numbers of moles can be represented as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} = K$$

Where, V_1 and n_1 are the initial volume and number of moles of the gas and V_2 and n_2 are the final volume and number of moles.

In calculations, we use Avogadro's law in a manner similar to the other gas laws.

Conceptual Check Point:

- **What will be your observation when?**
  i) The pressure inside the balloon is greater than external pressure.
  ii) The pressure inside the balloon is less than external pressure.
  iii) The pressure inside the balloon becomes equal to external pressure.
- **Why children don't want to inflate the balloon to a maximum extent?**

Example 4.4

Nitrogen peroxide gas contains 0.1 moles in 2.24 dm^3 balloon at STP. How many moles of this gas are present when the volume of gas becomes 11.2 dm^3 at same temperature and pressure?

Solution:

$n_1 = 0.1$ mol, $\quad V_1 = 2.24\ dm^3$

$n_2 = ?$ $\quad V_2 = 11.2 dm^3$

According to Avogadro'slaw,

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{or} \quad n_2 = \frac{n_1}{V_1} . V_2$$

By putting the values,

$$n_2 = \frac{0.1\text{mol}}{2.24\text{dm}^3} \times 11.2\,\text{dm}^3$$

$$n_2 = 0.5\text{mol}$$

Practice Exercise 4:

At constant temperature and pressure 10 moles of ammonia gas occupies a 224000mL volume in a closed container with a moveable piston. What will be the new volume of the gas when 2.5 moles of gas is removed from the container?

1.5 Ideal Gas Equation

Ideal gas equation is a relationship between pressure, volume, temperature and number of moles (amount) of a gas without assuming that any of the parameters is constant.

1.5.1 Derivation of Ideal Gas Equation

Ideal gas equation is derived by combining Boyle's law, Charles law and Avogadro's law.

According to Boyle's law:

$$V \propto \frac{1}{P} \qquad \text{(at constant } n \text{ and } T) \quad \ldots\ldots \text{(i)}$$

According to Charles law:

$$V \propto T \qquad \text{(at constant } n \text{ and } P) \quad \ldots\ldots \text{(ii)}$$

According to Avogadro's law,

$$V \propto n \qquad \text{(at constant } T \text{ and } P) \quad \ldots\ldots \text{(iii)}$$

By combining the above three equations, we have

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

This equation indicates that the volume of a gas is directly proportional to the number of moles of gas and to the temperature of the gas, but is inversely proportional to the pressure of the gas. We can replace the proportionality sign with equals sign by incorporating R.

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

By rearranging the above equation we get,

$$PV = nRT \quad \text{.......................} \quad (iv)$$

Where P is pressure, V is volume, n is number of moles, R is general gas constant and T is temperature.

Equation (iv) is called ideal gas equation and R is called general or universal gas constant. This equation is also called general gas equation or perfect gas law. The Ideal gas equation can be used to calculate the value of any one of the four variables P, V, T, and n, when the values of the other three variables are given.

General gas equation can be used to predict what happens when the pressure, volume, and temperature of a gas sample are all changed at once. We find that the value of pressure times volume divided by the Kelvin temperature is the same for the gas sample before and after the change. Hence, for one mole of gas, equation (iv) can be written as,

$$PV = RT \quad \text{or}$$

$$\frac{PV}{T} = R$$

If P, V and T are changed for a gas from P_1, V_1 and T_1 to P_2, V_2 and T_2, the equation will become:

$$\frac{P_1V_1}{T_1} = R$$

$$\frac{P_2V_2}{T_2} = R$$

Therefore,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{.......................} \quad (v)$$

This is another form of an ideal gas equation. This equation can be solved for any one of the six variables if the other five are known and is useful in dealing with the pressure-volume-temperature relationships of gases.

However, each of the quantities in the ideal gas equation must be expressed in the units within R.

Pressure (P) must be expressed in atmospheres, volume (V) in liters or decimeters cube, amount of gas (n) in moles and temperature (T) in Kelvins.

Conceptual Check Point:

What would have to be done to maintain a constant volume if the temperature increased?

Applications of Ideal Gas Equation

i) Ideal gas equation is used to calculate the mass of the gas, if P, V, T and molecular mass of the gas is known.

Since, $$n = \frac{\text{Mass of gas(m)}}{\text{Molecular mass of gas(M)}}$$

By putting the value of *n* in PV = nRT, we have

$$PV = \frac{m}{M}RT \qquad \text{or}$$

$$m = \frac{PVM}{RT} \qquad \text{.............................. (vi)}$$

ii) Ideal gas equation is used to calculate the molecular mass of the gas.
By re-arranging the equation (vi), we have

$$M = \frac{mRT}{PV}$$

Since, $\frac{M}{V}$ = density

Therefore,

$$M = d\frac{RT}{P} \qquad \text{.............................. (vii)}$$

iii) Ideal gas equation can be used to calculate density of the gas.
By re-arranging the equation (vii), we have,

$$d = \frac{PM}{RT}$$

The density of a gas is directly proportional to pressure and molar mass but inversely proportional to temperature. The higher the molar mass and pressure, the denser the gas. The higher the temperature, the less dense the gas. When we have equal molar masses of two gases at the same pressure but different temperatures, the hotter gas is less dense than the cooler one.

Conceptual Check Point:

Have you ever noticed that the tire pressure increases when the car has been recently driven?

4.5.2 Gas Constants and its Units

Ideal gas equation is used to determine the value of R. The value of R for one mole of a gas at STP is calculated as:

$P = 1atm$

$V = 22.414dm^3$

$n = 1mol$

$T = 273.15K$

$R = ?$

Formula used:

$$PV = nRT \quad \text{or} \quad R = \frac{PV}{nT}$$

By putting the values in the equation, we have

$$R = \frac{1atm \times 22.414dm^3}{1mole \times 273.15K}$$

$$= 0.08206 atm dm^3 mol^{-1} K^{-1}$$

If pressure is expressed in Newton per meter square (Nm^{-2}) and volume in cubic meter (m^3), then we find that,

$P = 101325Nm^{-2}$

$V = 0.0224m^3$

$n = 1mol$

$T = 273.15K$

$R = ?$

Formula used:

$$PV = nRT \quad \text{or} \quad R = \frac{PV}{nT}$$

By putting the values, we have

$$R = \frac{101325Nm^{-2} \times 0.022414m^3}{1mole \times 273.15K}$$

$$= 8.31 Nm\, mol^{-1} K^{-1}$$

Since,

$1\, Nm = 1\, J$

Therefore,

$R = 8.31\, J mol^{-1} K^{-1}$

Since,

$1\, cal = 4.184\, J$

Therefore,

$$R = \frac{8.31}{4.184}$$

$$= 1.986 cal\, mol^{-1} K^{-1}$$

Ideal gas equation can also be used to calculate the volume occupied by one mole of an ideal gas at STP.

$P = 1\text{ atm}$

$n = 1\text{ mole}$

$R = 0.08206\,dm^3\,atm\,mol^{-1}K^{-1}$

$T = 273.15\,K$

$V = ?$

Formula used:

$$PV = nRT \quad \text{or} \quad V = \frac{nRT}{P}$$

By putting the values, we have

$$V = \frac{1mole \times 0.08206 dm^3 atm\, mol^{-1} K^{-1} \times 273.15K}{1atm}$$

$$= 22.414\ dm^3$$

Example 4.5

At STP, the volume of carbon monoxide gas is 7.25 litre. What would be the volume of this gas at 29.4 psi and 25°C?

Solution:

Standard temperature $= T_1 = 273K$

Standard pressure $= P_1 = 1atm$

Initial volume of CO gas $= V_1 = 7.25L$

Final volume of CO gas $= V_2 = ?$

Final pressure $= P_2 = \dfrac{29.4psi}{14.7psi\ atm^{-1}} = 2atm$

Final temperature $= T_2 = 25°C + 273 = 298K$

According to combined gas law,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{or} \quad V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2}$$

By putting the values,

$$V_2 = \frac{(1atm)(7.25\ L)}{273\ K} \times \frac{298\ K}{2\ atm}$$

$$V_2 = 3.96L$$

Practice Exercise 5:

A sample of H_2S gas has a volume of $1.7dm^3$ at a temperature of 37°C. If gas has 3 moles, what will be the pressure of the gas?

Example 4.6

What is the molar mass of 134 grams of a gas at $-73^\circ C$? The pressure of gas is 10 atm and its volume is 5 dm^3.

Solution:

Molar mass $= M = ?$

Mass of gas $= m = 134g$

Temperature $= T = -73^\circ C + 273 = 200K$

Pressure $= P = 10atm$

Volume $= V = 5dm^3$

Formula used: $M = \dfrac{mRT}{PV}$

By putting the values,

$$M = \frac{(134g)(0.08206atm\ dm^3mol^{-1}K^{-1})(200K)}{(10atm)(5dm^3)}$$

$$M = 44g\ mol^{-1}$$

Practice Exercise 6:

A syringe has 20mL of gas at 47°C and 50.6625kPa pressures. What is the molar mass of the gas in Kilograms? The mass of gas is 1.52 grams.

Example 4.7

Calculate the density of H_2S gas in g/dm^3 at $0^\circ C$ and 960mmHg pressure.

Solution:

Density of H_2S gas $= d = ?$

Temperature $= T = 0^\circ C + 273 = 273K$

Pressure $= P = 960mmHg / 760mmHg\ atm^{-1} = 1.26atm$

Molar mass of H_2S gas $= M = 34g\ mol^{-1}$

Formula used: $d = \dfrac{PM}{RT}$

By putting the values,

$$d = \frac{(1.26\text{atm})(34\text{gmol}^{-1})}{(0.08206 \text{ atm dm}^3\text{mol}^{-1}\text{K}^{-1})(273\text{K})}$$

$$d = 1.91\text{g mol}^{-1}$$

Practice Exercise 7:

What is the density in g/dm^3 of butane (C_4H_{10}) gas at 100°C and one atmosphere pressure?

4.6 Deviation of Real Gases from Ideal Gas Behaviour

The gases which obey gas laws strictly at all conditions of temperatures and pressures are called ideal or perfect gases while those which do not obey gas laws strictly are known as non-ideal or real or actual gases. The behavior of a real gas is often somewhat different from that of an ideal gas. In actual practice no gas is ideal. All known gases are non-ideal. The gases at sufficiently high temperature and low pressure behave ideally.

4.6.1 Graphical Explanation

A useful measure of how much a gas deviates from ideal gas behavior is found in its compressibility factor. The compressibility factor (Z) of a gas is the ratio PV/RT. From the ideal gas equation (for one mole of a gas) we see that for an ideal gas, the ratio PV/RT = Z = 1. The compressibility factor values can be less than or more than one for a real gas. When the pressure is increased for real gas, the volume of one mole of the gas is decreased because of strong intermolecular forces. As a result of this, the value of compressibility factor (Z) becomes less than one ($Z < 1$). When the pressure is further increased (i.e. at very high pressure), the gas molecules are crowded closely together and the intermolecular repulsions become operative, and by doubling the pressure cannot halve the total volume. As a result, the actual volume of a real gas is larger than expected for an ideal gas, and the value of compressibility factor (Z) becomes greater than one ($Z > 1$). The ratio PV/RT approaches 1 only at very low pressures (≤ 10 atm.) for real gases.

The extent of deviation depends on the nature of the gas, temperature and pressure under which the behaviour of a real gas is studied. The deviation (departure) from ideal behaviour is most significant at *high pressures* and *low temperatures*, that is, near the conditions under which the gas liquefies.

) Deviation at high pressure

In order to check the ideality of a gas we plot a graph between pressure on X-axis and compressibility factor (Z) on y-axis for four real gases and an ideal gas at a given temperature. In case of ideal gas a straight line parallel to x-axis is obtained. But all real gases (H_2, He, N_2 and NH_3) has been found to show marked deviation from ideal behaviour.

At one atmospheric pressure (and room temperature), most gases behave nearly like an ideal gas, because

i) The molecules are moving so rapidly and are so far apart that the attractions between them are very weak or negligible.

ii) The space between the molecules is so large that the volume occupied by the molecules themselves is insignificant (negligible).

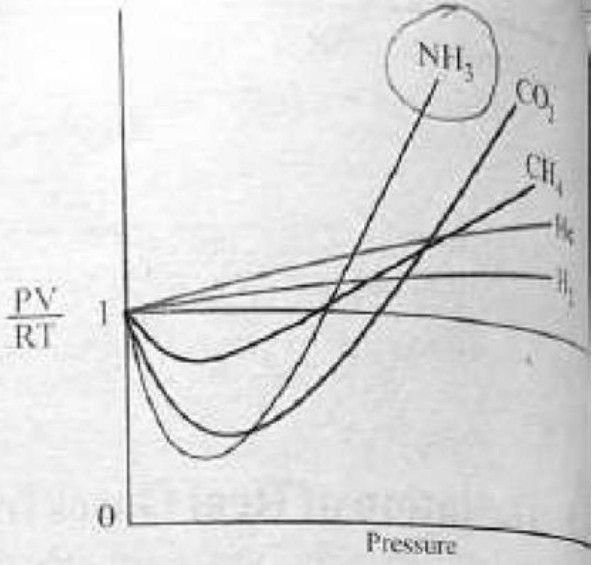

Figure 4.8: It is clear from the graph t deviation of gases from the id behaviour is greater at high pressure. Th is due to more attractive forces at ro temperature.

At high pressure, the gas volume decreases and the empty spaces among g molecules decrease. As a result of this, the volume of actual gas molecules beco significant (does not remain very small as compared to the volume of containe total volume of the gas) and hence cannot be neglected.

Different gases deviate differently from ideal behaviour. The deviatio real gases from ideal behaviour is higher for polar gases and lower for non-p gases. For gases such as hydrogen, helium, nitrogen, neon, or oxygen, devia from the ideal gas behaviour are less than 0.1 percent at room temperature and atmosphere pressure. Other gases, such as CO_2, SO_2, NH_3, or H_2O, have stro intermolecular forces and hence show greater deviation from ideal behaviour.

ii) Deviation at Low Temperature

As the temperature of gas molecules decreases, the kinetic energy and velocit gas molecules decreases. As a result of this, the number of collisions decreases the gas molecules come close to each other. thus forces of attractions among molecules become effective. Hence at very low temperature, near the tempera under which the gas liquefies, the gases deviate from ideal behaviour.

4.6.2 Causes for Deviation

The deviation is due to two faulty assumptions:

i) The actual volume of the gas is very small as compared to the volume of the gas or the volume of the container.

ii) There is no force of attraction among gas molecules.

The postulate (i) is responsible for the deviation at very high pressure and postulate (ii) is responsible for the deviation at very low temperature..

4.7 Van der Waal's Equation

Johannes Van der Waals (1837-1923)

The Dutch physicist Johannes Van der Waals, in 1873, at the University of Amsterdam modified the general gas equation in order to calculate the volume and pressure of a real gas. He recognized that the ideal-gas equation could be corrected by performing the volume correction and pressure correction. He got a Nobel Prize in 1910 for his work.

i) Volume Correction

When an actual gas is compressed, the molecules come close to each other and a stage comes when it is not possible to compress it any more. At this stage, they produce force of repulsion, because gas molecules have definite volume. The actual volume of gas molecules is very small but it cannot be neglected. If the effective volume of the molecules per mole of a gas is b, then the available (free) volume to gas molecules is the volume of the container minus the volume of gas molecules.

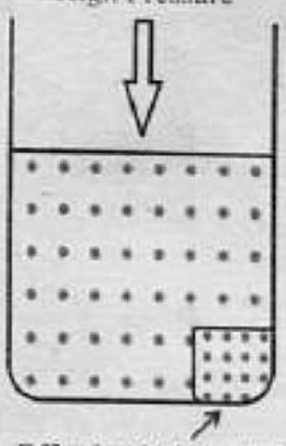

Figure 4.9

$$V_{free} = V_{container} - V_{molecules}$$

or

$$V_{free} = V_{container} - b$$

This equation can be written as:

$$V_{real} = V_{ideal} - nb \quad \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ (i)$$

P′ will be proportional to the number of molecules per unit volume colliding with the walls of container and also to the number of molecules per unit volume dragging them.

Where, 'n' is the number of moles and b' is effective or excluded volume of gas. The volume occupied by one mole of a gas in highly compressed state is called effective volume. It is four times of the actual volume of the gas molecules.

$$b = 4\,Vm$$

Where, Vm is the actual volume of one mole of a gas.

ii) Pressure Correction

Consider a molecule in the interior of a gas is attracted by other molecules from all sides and there is no net attraction on the molecule. When such molecules strike with the walls of container, it cannot exert full pressure due to backward pull of gas molecules. So these molecules do not collide with a force with which it should collide. Hence less pressure is exerted on the wall than actual pressure (ideal pressure).

$$P_{obs} = P_i - P'$$

or $$P_i = P_{obs} + P' \quad \text{..................} \quad (ii)$$

Where,

P_{obs} is observed pressure, P_i is ideal gas pressure and P' is decrease in ideal pressure

The P' will decrease as the volume will increase due to decrease in forces of attractions. Mathematically,

$$P' \propto \frac{n}{V} \times \frac{n}{V}$$

or

$$P' \propto \frac{n^2}{V^2}$$

or

$$P' = a\frac{n^2}{V^2}$$

Where, "**a**" is proportionality constant.

By putting the value of **P'** in equation (ii), we have

$$P_i = P_{obs} + \frac{an^2}{V^2} \quad \text{..........................} \quad (iii)$$

By putting the value of equation (i) and equation (iii) in the general gas equation (PV=RT), we have

$$(P_{obs} + \frac{a\,n^2}{V^2})(V - n\,b) = nRT$$

This is Vander Waal's equation where *'a'* and *'b'* are experimentally determined constants called Vander Waal's constants. When 'a' and 'b' are both the van der Waal's equation reduces to the ideal gas equation. In this equation, P, V, T, and n represent the measured values of pressure, volume, temperature, and number of moles, respectively, just as in the ideal gas equation.

The values of constants 'a' and 'b' for real gases are given in the table.

Table 4.3: Van der Waal's Constants for Some Real Gases:

Name	Formula	a (atm dm^{6}mol^{-2})	b (dm^{3}mol^{-1})
Hydrogen	H_2	0.244	0.0266
Helium	He	0.034	0.0237
Nitrogen	N_2	1.390	0.0391
Oxygen	O_2	1.360	0.0318
Argon	Ar	1.345	0.0322
Krypton	Kr	2.318	0.0398
Methane	CH_4	2.253.	0.0428
Ammonia	NH_3	4.170	0.0371
Water	H_2O	5.464	0.0305
Carbon dioxide	CO_2	3.592	0.0427
Hydrogen chloride	HCl	3.667	0.0408

It is clear from the above table that polar gases have greater values of 'a' than non-polar gases. This is due to the presence of strong intermolecular forces in polar gases as compared to non-polar gases.

4.8 Dalton's Law of Partial Pressure

The English scientist John Dalton (1766-1844), in 1803, made an important contribution to the study of gaseous mixtures. Dalton's law of partial pressures states that the total pressure exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of all the gases present in the mixture.

Suppose in a system three gases A, B, and C are present. The partial pressures of these gases are:

P_A = Partial pressure of gas A

P_B = Partial pressure of gas B

P_C = Partial pressure of gas C

Then Dalton's law may be mathematically written as:

$$P_T = P_A + P_B + P_C + \ldots$$

Where, P_T is the total pressure of mixture of gases.

The air we breathe is a mixture of about 78.084% nitrogen, 20.948% oxygen, 0.934% argon and, 0.033% carbon dioxide and other gases (Hydrogen, helium, methane, krypton, carbon monoxide, xenon, ozone, ammonia, nitrogen dioxide, suphur dioxide and water vapours) in trace amounts. Due to the force of gravity, molecules making up the atmosphere are most concentrated near Earth's surface. In dry air, total pressure is the sum of partial pressure of N_2, O_2, Ar, CO_2 and other gases.

Interesting information:

The composition of air do[es] not change appreciably wi[th] altitude, but the tot[al] pressure decreases rapid[ly]. The partial pressure o[f] oxygen in air therefo[re] decreases with increasin[g] altitude, and it is this chang[e] that leads to difficulty i[n] breathing at high elevation[s].

Mathematically, it can be expressed as,

$$P_{atm} = P_{N_2} + P_{O_2} + P_{Ar} + P_{CO_2} + P_{other\,gases}$$

Suppose the partial pressure of nitrogen, oxygen, argon, carbon dioxide, an[d] other gases in the air is 593 torr, 159.5 torr, 7.14 torr, 0.23 torr, and 0.13 to[rr] respectively. So the total pressure of the gases in the air will be:

$$P_{atm} = 593 + 159.5 + 7.14 + 0.23 + 0.13$$

$$P_{atm} = 760\text{torr}$$

A partial pressure is the pressure of an individual gas in the mixture of gase[s]. This law says that the pressure of a gas remains same whether gas is alóne or mix[ed] with other gases at constant temperature and pressure.

Put another way, pressure of a gas is due to collision of molecules. It depen[ds] upon number of molecules or moles of a gas present in the mixture. Greater t[he] number of moles (molecules) in a mixture, greater would be the pressure. It means

$$P \propto n$$

So, pressure of a gas is directly proportional to the number of moles of a ga[s]. For instance, if we have 1000 gaseous molecules at 50°C in a container and th[ey] exert a pressure of 0.1atm, then adding another 1000 gaseous molecules of [a] substance at the same temperature to the container would increase the pressure [to] 0.2atm.

Conceptual Check Point:

Imagine what would happen if someone starts filling the air in the tyre of a bike a[nd] cannot stop this process?

Expressions for Calculation of Partial Pressure of a Gas

i) From general gas equation

In a mixture of gases, the partial pressures of gases can be calculated from [the] general gas equation when number of moles and total pressure of each gas is g[iven] at constant temperature and volume.

Suppose two gases A and B are present in the gas cylinder having partial pressures P_A and P_B respectively. The number of moles of gases A and B are n_A and n_B. The general gas equation is $PV = nRT$. Since R, T and V are same for gases A and B, hence the partial pressures of these gases are:

$$P_A = n_A \frac{RT}{V} \quad \text{..........................} \quad \text{(i)}$$

Similarly,

$$P_B = n_B \frac{RT}{V} \quad \text{..........................} \quad \text{(ii)}$$

For a mixture of two gases A and B, the total pressure is given by the expression:

$$P_{Total} = P_A + P_B = (n_A + n_B)\frac{RT}{V}$$

or $$P_{Total} = n_{Total} \cdot \frac{RT}{V} \quad \text{..........................} \quad \text{(iii)}$$

Conceptual Check Point:

Suppose we have 2 separate tanks, one filled with nitrogen gas (N_2) at 450 mmHg and the other with oxygen at 250 mmHg. When both of the gases are transferred to a single tank with the same volume and temperature, then, what would be the total pressure of the gas mixture in this tank?

By dividing equation (i) by (iii) we get,

$$\frac{P_A}{P_T} = \frac{n_A \frac{RT}{V}}{n_T \frac{RT}{V}}$$

$$\frac{P_A}{P_T} = \frac{n_A}{n_T}$$

$$P_A = \frac{n_A}{n_T} \cdot P_T \quad \text{..........................} \quad \text{(iv)}$$

here,

P_A is partial pressure of gas A, n_A is number of moles of gas A, n_T is total nber of moles of gases and P_T is total pressure of gases

nilarly,

$$P_B = \frac{n_B}{n_T} P_T$$

ii) From Mole Fraction of Gases

The number of moles of a particular gas divided by total number of moles of ga present in the mixture is called mole fraction and is represented by X. therefore,

$$\frac{n_A}{n_T} = X_A \qquad (n_T = n_A + n_B)$$

By putting the value of n_A / n_T in equation (iv) we have,

$$P_A = X_A . P_T$$

Where, X_A is mole fraction of gas A. This equation tells us that the pressu of a gas in a mixture of gases is the product of its mole fraction and the total pressu of the mixture. You can determine the partial pressure of gas A (P_A) by multiplyin its mole fraction with the total pressure (P_T).

Similarly

$$P_B = X_B . P_T$$

Where, X_B is mole fraction of gas B.

For example, the mole fraction of oxygen in air is 0.21; hence its par pressure is 0.21 atm or 159.5 torr at standard temperature and pressure.

Applications of Dalton's Law

i) Determining the pressure of pure and dry gases

Some gases like H_2, N_2 and O_2 are collected by downward displacement of wa Due to evaporation, some water molecules are mixed with gas and gas beco moist. The partial pressure of such gases can easily be determined by using equation of Dalton's law of partial pressure:

$$P_{gas} = P_T - P_{water vapours}$$

$$P_{H_2} = P_T - P_{water\ vapours}$$

The partial pressure of water vapours in gases is called aqu tension or vapour pressure of water.

Table 4.4: Vapour Pressure of Water at Various Temperatures:

Temperature (°C)	Vapour Pressure (mmHg)	Temperature (°C)	Vapour Pre (mmHg)
0	4.6	60	149.4
10	9.2	70	233.7
20	17.5	80	355.1
30	31.8	90	525.8
40	55.3	100	760.0
50	92.5		

ii) **Process of Respiration**

Respiration is due to difference in partial pressure of O_2 in air (159 torr) and lungs (116 torr). Oxygen flows from a region of higher partial pressure to a region of lower partial pressure. The partial pressure of CO_2 inside the body is greater than in atmosphere. So CO_2 comes out from the lungs.

iii) **At higher Altitude**

At higher altitude, pilot feels uncomfortable where pressure is around 150 torr which is less than 159 torr (at which we feel comfortable). So pilot uses pressurized cabins. On top of the highest mountain, Mt. Everest, the total atmospheric pressure is 270 torr, so the partial pressure of oxygen is only 57 torr, or about one-third of normal. That is why a human cannot survive for long at such a low pressure of oxygen. At that altitude, the mountaineers use an oxygen tank and mask, which give an increased partial pressure of oxygen to the lungs.

iv) **Deep Sea Divers**

The partial pressure of air (O_2) increases at the depth below the surface of water. Partial pressure of oxygen increases five times at the depth of 40 meters. Ordinary air cannot be used there. So, deep sea divers use a mixture of helium (96%) and oxygen (4%) called ***heliox*** that contains a lower percentage of O_2 than air. Helium gas does not dissolve well in blood, and thus it is safer for a diver to inhale this oxygen-helium mixture. At the same time, the oxygen exerts the same pressure that it would normally at one atmosphere pressure.

Conceptual check point:

- Explain why the pressure goes up when gas molecules are added to a sample of gas in a fixed-volume container at constant temperature?
- What happens to the partial pressure of each gas in a mixture when the volume is decreased by (a) lowering the temperature or (b) increasing the total pressure?

Example 4.8

A tube light which has a volume of 500 mL contains 10 grams of argon and 0.5 grams of neon at room temperature (25°C). Calculate the total pressure of the mixture of gases in mmHg.

Solution:

Volume of gases $= V = \dfrac{500\text{mL}}{1000\text{mL/L}} = 0.5\text{L}$

Mass of argon $= 10\text{g}$

Mass of neon $= 0.5\text{g}$

Temperature $= T = 25^{\circ}C + 273 = 298K$

Total pressure $= P_T = ?$

Before going to calculate total pressure,

We should know about moles of gases.

$$\text{Moles of argon} = n_{Ar} = \frac{10g}{39.948 g mol^{-1}} = 0.250 mol$$

$$\text{Moles of neon} = n_{Ne} = \frac{0.50g}{20.179 g mol^{-1}} = 0.025 mol$$

According to Dalton's law,

$$P_T = (n_{Ar} + n_{Ne})\frac{RT}{V}$$

By putting the values,

$$P_T = (0.250\ mol + 0.025 mol)\frac{(0.08206 atm\ dm^3 mol^{-1} K^{-1})(298K)}{(0.50K)}$$

$$P_T = 13.45 atm$$

The total pressure of gases in mmHg can be calculated as,

$$P_T = 13.45\ atm \times 760\ mmHg\ atm^{-1} = 10222\ mmHg$$

Practice Exercise 8:

What is the total pressure of 3.50 grams of oxygen and 9.25 grams of nitrog gases at 9°C ? Both of the gases are present in 100 mL perfume bottle.

4.9 Graham's Law of Diffusion and Effusion

Diffusion is the spontaneous mixing of gas molecules by random motion where molecular collision occurs. When a perfume bottle is opened at the front of a classroom, it takes some time before everyone in the room can smell it, because time is required for the perfume molecules to mix with the air. Lighter molecules diffuse more quickly than heavier ones, so the first molecules you would smell from a perfume mixture are the lighter ones. The rate of diffusion is the rate of the mixing of gases. Effusion is the escape of gas molecules through a

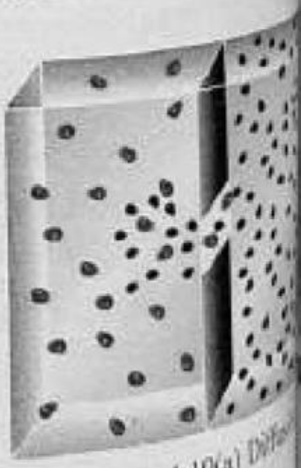

Figure 4:10(a) Diff

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tiny (small) hole into a vacuum without molecular collision. Lighter molecules effuse more quickly than heavier ones. The rate of effusion is inversely proportional to molar mass of the gas. If we have two balloons in which one has helium and the other has carbon dioxide gas. A helium filled balloon deflates rapidly as compared to carbon dioxide filled balloon because the rate of effusion through the tiny holes of the balloon is faster for the lighter helium atoms than the heavier carbon dioxide molecules.

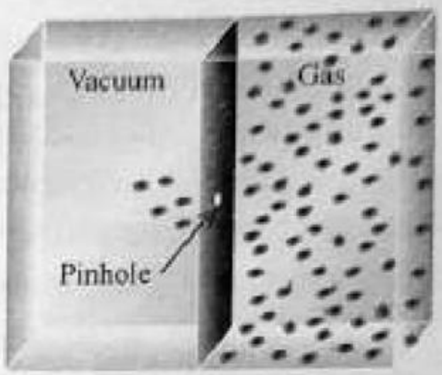

Figure 4.10(b): Effusion

The rates at which both of these processes occur depend on the speeds of gas molecules; the faster the molecules move, the more rapidly diffusion and effusion occur. In 1846, Thomas Graham, a Scottish chemist, studied the rates of diffusion and effusion of gases. He recognized that the rate of diffusion and effusion of a gas is related to their molar masses or densities.

This relation is known as Graham's Law of diffusion and effusion, which is stated as,

"The rate of diffusion or effusion of two gases is inversely proportional to the square roots of their densities or molar masses at constant temperature and pressure".


Thomas Graham
(1805-1869)

Graham's law is usually used to compare the rates of diffusion and effusion of different gases, so the proportionality constant can be eliminated and an equation can be formed by writing the ratio of diffusion and effusion rates:

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

or

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

Where,

r_1 is rate of diffusion or effusion of gas 1, r_2 is rate of diffusion or effusion of gas 2, d_1 is density of gas 1, d_2 is density of gas 2, M_1 is Molecular mass of gas 1 and

M_2 is Molecular mass of gas 2.

Gases with low molar masses diffuse (and effuse) more rapidly than gas with high molar masses. Thus, H_2 with a molar mass of 2 will diffuse more rapidly than ammonia, NH_3, with a molar mass of 17.

The rate of effusion of molecules from a container depends on three factors:

i) The cross-sectional area of the hole (the larger the hole, the more likely the gas molecules are to pass through the hole);

ii) The number of molecules per unit volume (the more congested the molecules are, the more likely they are to come across the hole); and

iii) The average molecular speed (the faster the gas molecules are moving the more likely the gas molecules are to escape).

Demonstration of Graham's Law

Take 100 cm long glass tube open at both ends. Put two cotton plugs soaked in HCl and NH_3 solutions at each end of the tube simultaneously. The vapours of NH_3 and HCl escape from their respective ends and producing a white ring of NH_4Cl where they meet several minutes later. The distance travelled by ammonia is 59.5 cm and by HCl gas is 40.5 cm.

$$NH_{3(g)} + HCl_{(g)} \longrightarrow NH_4Cl_{(s)}$$

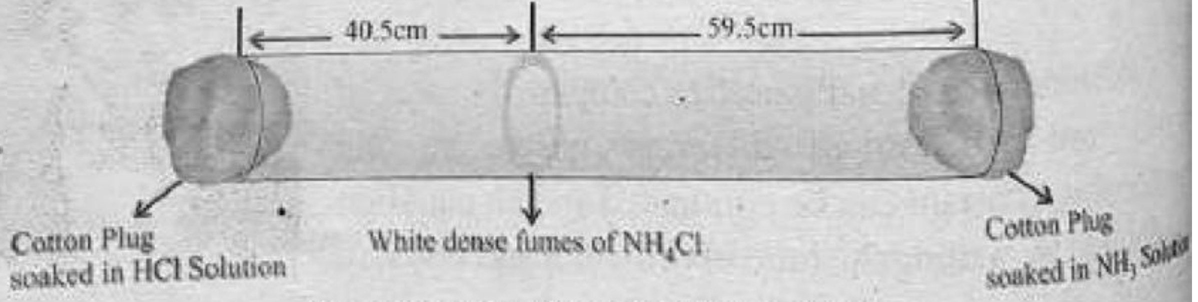

Figure 4.11: Verification of Graham's Law

The ratio of velocities (rates) of two gases depends on their distance traveled by the different gases at the same time.

$$\frac{\text{The rate of diffusion of } NH_3}{\text{The rate of diffusion of HCl}} = \frac{\text{Distance traveled by } NH_3}{\text{Distance traveled by HCl}} = \frac{59.5\text{cm}}{40.5\text{cm}} = 1$$

This means that the rate of diffusion of NH_3 gas is 1.46 times greater than HCl gas, because NH_3 gas is lighter than HCl.

Conceptual Check Point:

Why does the hydrogen leak out from the balloon at a higher rate than the nitrogen?

Example 4.9

An unknown gas (hydrocarbon) diffuses two times as fast as SO_2 gas at constant temperature. Calculate the molar mass of unknown gas and suggest what this gas might be?

Solution:

Molar mass of SO_2 gas $= M_{SO_2} = 64 gmol^{-1}$

Molar mass of unknown gas $= M_x = ?$

Rate of diffusion of SO_2 gas $= r_{SO_2} = 1$

Rate of diffusion of unknown gas $= r_x = 2$

According to Graham's law,

$$\frac{r_x}{r_{SO_2}} = \sqrt{\frac{M_{SO_2}}{M_x}}$$

By putting the values,

$$\frac{2}{1} = \sqrt{\frac{64 gmol^{-1}}{M_x}}$$

By taking square on both sides,

$$\frac{4}{1} = \frac{64 gmol^{-1}}{M_x} \quad \text{or}$$

$$M_x = \frac{64 gmol^{-1}}{4} = 16 g\,mol^{-1}$$

The unknown gas is methane, because it has hydrogen and carbon atoms and has molar mass $16 gmol^{-1}$.

Practice Exercise 9:
Calculate the rate of diffusion of equal volumes of NH_3 and HCl at constant temperature.

4.10 Liquefaction of Gases

The conversion of gases into liquids is called liquefaction of gases. The conversion of gas into liquid needs high pressure and low temperature.

At high pressure, molecules of a gas come close to each other and at low temperature, attractive forces among gas molecules increase.

Keep in mind, the gaseous substance is capable of being liquefied by compressing when it has a temperature below its critical temperature. Criti temperature is the temperature above which a gas cannot be liquefied no matter h much pressure is applied. It is denoted by T_C. If room temperature is *below criti temperature*, the gas can be liquefied by applying sufficient pressure. If ro temperature is above critical temperature, then the gas can only be liquefied at t cost of added pressure and a lowering of temperature to a value below criti temperature. For example, carbon dioxide is a gas at room tempera (approximately 25°C). This is below its critical temperature of 31.1°C. If the CO is gradually compressed, a pressure will eventually be reached that will cause CO_2 to liquefy. Gases such as O_2 and N_2, which have critical temperatures far bel 0°C, can never be liquefied at room temperature. When they are compressed, t simply become high pressure gases. To make liquid N_2 or O_2, the gases must made very cold as well as be compressed to high pressures.

The intermolecular attraction is a finite quantity for any given substa Below T_C, this force is sufficiently strong to hold the molecules together (und some appropriate pressure) in a liquid. Above T_C, molecular motion becomes energetic that the molecules can always break away from this attraction. Criti temperature depends upon size and shape of gas molecules and the strength intermolecular forces. The pressure required to liquefy the gas at its criti temperature is called critical pressure. It is denoted by P_C. The volume occupied one mole of a gas at critical temperature and critical pressure is called criti volume. It is denoted by V_C.

Table 4.5: Critical Temperature, Pressure and Volume of Some Gases

Substance	Formula	T_c(°C)	P_c(atm)	V_c($cm^3\ mol^{-1}$)
Helium	He	−267.9	2.26	62
Hydrogen	H_2	−239.9	12.8	70
Nitrogen	N_2	−147.1	33.5	90
Carbon monoxide	CO	−138.8	35.0	90
Argon	Ar	−122.2	48.0	78.5
Oxygen	O_2	−118.8	49.0	74
Methane	CH_4	−83.1	45.6	99
Carbon dioxide	CO_2	31.1	73.0	94
Ammonia	NH_3	132.5	111.5	72
Chlorine	Cl_2	133.9	76.1	123.8
Sulphur dioxide	SO_2	157.1	77.7	123
Methyl alcohol	CH_3OH	239.9	78.5	118
Water	H_2O	374.0	217.7	56

The gases can be liquified by Joule-Thomson effect. According to Joule Thomson effect "When highly compressed gases are suddenly allowed to expand, they produce cooling and the temperature of gas falls to such an extent that it changes into liquid state."

The process is called Joule-Thomson or Joule-Kelvin effect.

The molecules of the compressed gas are very close to each other and have appreciable attractive forces. When this gas is allowed to expand suddenly through the nozzle of jet, then the molecules move apart from each other. For expansion, some amount of energy is required to break intermolecular forces. This energy is taken from the gas itself, therefore the temperature of gas falls and cooling is produced.

Lind's Method

In 1895, Lind used the principle of Joule-Thomson effect for the liquefaction of air (O_2 and N_2). Lind's Method is used for liquefaction of all gases except H_2 (T_C=33.3 K) and He (T_C = 5.3 K) which have critical temperature values very close to absolute zero. The apparatus used for this process is: (i) Compressor (ii) Cooler (iii) Expansion Chamber (iv) Spiral pipe

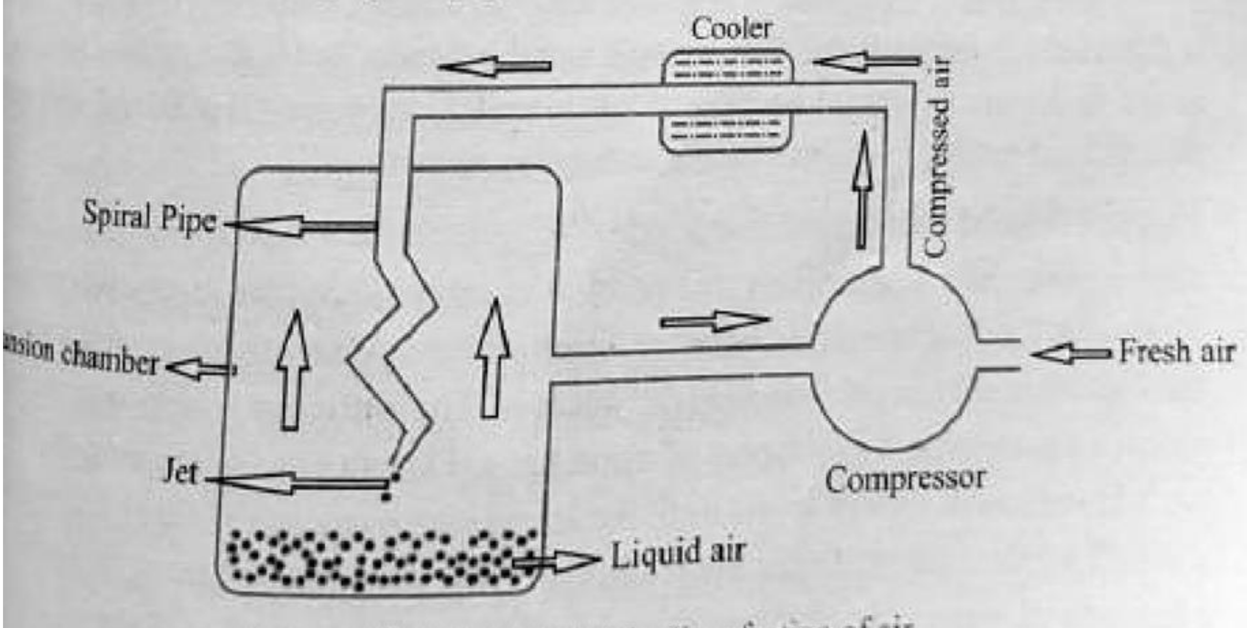

Figure 4.12: Lind's method for liquefaction of air

For the liquefaction of air, pure dry air is compressed to about 200 atm by a pressor. Some heat is produced due to compression. This heat is removed by ing gas through water cooled pipe or cooler. Compressed air is then passed gh a spiral pipe having a jet at the end. When the air comes out suddenly from et, the expansion takes place from 200 atm to 1atm. In this way, the temperature of gas falls and it is cooled. This cooled air goes up and cools the incoming compressed air and then returns to the compressor. This process is repeated again and again till the air is completely liquefied.

Keep in Mind

Liquefied natural gas (LNG) and liquefied petroleum gas (LPG) are among the examples of liquefied gas in daily use. In both cases, the volume of the liquefied gas is far less than it would be if the gas were in a vaporized state, thus enabling ease and economy of transport. Liquefied gases have a lot of applications. They are used for heating purposes at homes and as a fuel for cars, boats etc. Liquefied hydrogen and oxygen gases are used as a fuel in rocket engines; and liquefied oxygen and petroleum is used in welding.

4.11 The Fourth State of Matter: Plasma

The fourth very high-temperature physical state of matter which has an ionized g mixture of ions, electrons and neutral atoms, is called plasma state.

It was first identified in a Crooks tube by British English chemist and physicist, William Crooke in 1879 and he called this state a "radiant matter". The w "PLASMA" was first applied to ionized gas by Dr. Irving Langmuir (1881-19 an American chemist and physicist, in 1929.

Plasma is often called the "fourth state of matter", assuming that solid, liq and gas are respectively the first, second and third state. But is should be consid as the first state of matter because it has been known since "Big Bang" and m than 99% of matter in the universe is in the plasma state.

Formation of Plasma

Matter on Earth has electrons that revolve in orbit around the nucleus in an at There is a force of attraction between electrons and positively charged nucleus the electrons stay in orbit around the nucleus. The sufficient energy is provid atoms to overcome these forces of attractions. Plasma comes into existence we add sufficient energy to gas until the atoms lose some or all of their electro a result of this, atoms break apart into positive ions and electrons, or some even into atomic nuclei and free electrons. This state is known as plasma. At at least partially ionized. The degree of ionization does not have to be too la the size of the plasma formation is big enough. Even a partially ionized gas in as little as 1% of the particles are ionized can have the characteristics of plasma response to magnetic fields and high electrical conductivity). The energy r to change gas into plasma can be of various origins: thermal, electrical, (ultraviolet light or intense visible light from a laser).

Properties of Plasma

Plasmas have their own unique qualities just as solids, liquids and gases do.

i) Plasma state is overall neutral because it has equal number of electrons and positive ions.

ii) Plasmas are good conductors of electricity and are affected by magnetic fields because they are composed of ions (negatively charged electrons and positively charged nuclei).

iii) Plasmas neither have definite shape nor definite volume like gases because the particles can move past one another.

iv) Plasmas are easily compressible because their particles are far apart from each other.

v) The free electrons in a metal are considered as plasma.

Comparison of Plasma and Gas Phases

The characteristics of plasmas are significantly different from those of ordinary neutral gases so that plasmas are considered a distinct "fourth state of matter."

It differs from gases in number of ways:

i) Plasma is different from a gas, because it is made up of groups of positively and negatively charged particles and are strongly influenced by electric and magnetic fields while neutral gases are not.

ii) Gases are excellent insulators due to absence of charged particles while plasmas are good conductors of electricity due to presence of charged particles.

Applications of Plasma

Plasmas occur naturally but also can be easily produced in a laboratory and in industry. Numerous technologies make use of plasma.

i) They light up our offices and homes, make our computers and electronic equipment work.

ii) They drive lasers, help clean up the environment, and make tools corrosion-resistant.

iii) It is found in fluorescent lamps to produce light. When electric current is passed through a tube containing neon gas at low pressure, the neon atoms absorb the energy and use it to excite electron from the (filled) second shell to a less stable empty shell lying much farther from the nucleus. The electron doesn't stay in the far shell for long. When it comes back from the higher energy empty shells to the proper lower energy shell, it releases the extra energy as visible light of various colours. The predominant colour of the light emitted by neon is orange-red. "Neon" lights having other colours use a different gas or colored glass. In particular, argon, krypton, and xenon are also used in some "neon" lights. Hydrogen gas glows with green colour, helium glows with pink colour, nitrogen glows with purple colour whereas argon glows with a blue colour. HG, HeP. NP. AB

iv) It can be used for cleaning and sterilization of food and operation theaters.

v) Plasma can be used to destroy bacteria, viruses, fungi and spores.

vi) Our computers and electronic devices work on the basis of plasma. You are well aware of plasma TV and cell phones which are manufactured using plasma.

vii) Most of the synthetic fibers used in clothing, and advanced packaging materials are plasma treated.

viii) It is used in automobile industry in painting the bumpers and other parts made up of plastic.

ix) Plasma is used for the treatment of exhaust gases and also for the treatment of burns, ulcers and other skin diseases.

x) Plasma torches are used in industry to cut metals.

Plasma is expected to soon be widely used in surgery, decontamination and sterilization of surfaces and devices, and air and water streams, as well as in direct treatment of skin diseases.

Summary of Facts and Concepts

- There are four main states of matter. Solids, liquids, gases, and plasmas. They are all different states of matter. Each of these states is also known as a phase. Substances can be changed from one phase to another by increasing or decreasing temperature.
- Gases exert pressure, which is the force per unit area. The SI unit of pressure is Pascal (Pa).
- The behavior of gas is described in terms of its pressure, temperature, volume and amount (number of moles).Equations relating these four parameters are called the gas laws.
- Boyle's law states that the volume of a sample of gas is inversely proportional to pressure at constant temperature.

$$V \propto \frac{1}{P} \qquad \text{(at constant temperature)}$$

- Charles's law states that the volume of a sample of gas is directly proportional to its absolute temperature at constant pressure.

  $V \propto T$ (at constant pressure)

- Avogadro's law states that the volume of a sample of gas is directly proportional to the number of moles at constant temperature and pressure.

  $V \propto n$ (at constant temperature and pressure)

- The general gas equation ($PV = nRT$) is obtained by combining the laws of Boyle, Charles, and Avogadro and relates pressure, volume, temperature, and number of moles without assuming that any of the parameters is constant.
- The temperature $0^{\circ}C$ and 1atm pressure is known as standard temperature and pressure (STP).
- Dalton's law of partial pressures states that the total pressure exerted by a mixture of non-reacting gases is the sum of the partial pressures of the components.

  $P_T = P_A + P_B + P_C + \ldots$

  This law is used to calculate the partial pressure of gases.

- The kinetic molecular theory of gases is a model that accounts for ideal gas behavior. The behavior of most gases is nearly ideal except at very high pressures and low temperatures.
- Diffusion is the mixing of two or more gases whereas effusion is the escape of gas molecules from a container through a small hole into an empty chamber.
- The departure of a real gas from ideal gas behavior is called deviation. Deviation from ideal gas behavior increase in magnitude as pressure increases and as temperature decreases.
- Van der Waal's equation is a modified equation which is used to calculate the pressure and volume of real gases.
- Gases can be liquefied by applying pressure at critical temperature or below it. Gases can't be liquefied above critical temperature.
- Plasma is considered as the fourth state of matter. It consists of ions, electrons and neutral particles. About 99% of the visible universe is made up of plasma.

Questions and Problems

Q.1. Four answers are given for each question. Select the correct one:

i) Matter exists in the number of states:
(a) Two (b) three (c) four (d) five

ii) The SI unit of pressure is:
(a) Atmosphere (b) Pascal
(c) Newton (d) Torr

iii) The value of general gas constant, R in SI system is:
(a) $8.31 NmK^{-1}mol^{-1}$ (b) $8.31 atm\, dm^3 mol^{-1} K^{-1}$
(c) $0.0821 Nm^{-1} mol^{-1}$ (d) $0.0821 atm\, dm^3 mol^{-1} K^{-1}$

iv) The instrument that is used to measure the atmosphere pressure is:
(a) Thermometer (b) Manometer
(c) Barometer (d) Voltmeter

v) When the Kelvin temperature of a gas is doubled, the volume will:
(a) Remain same (b) be doubled
(c) reduce to one half (d) increase four times

vi) When the Kelvin temperature of the gas is doubled and the pressure is reduced to one-half, then the volume of the gas will:
(a) Remain same (b) be doubled
(c) reduce to one-fourth (d) increase four times.

vii) Which one of the following gases diffuse more rapidly:
(a) N_2 (b) CH_4 (c) O_2 (d) CO_2

viii) All the gases can be liquefied by Lind's method except:
(a) H_2 and NH_3 (b) H_2 and N_2
(c) H_2 and He (d) N_2 and O_2

ix) What is the rate of effusion of an unknown gas of molar mass 50 compared to hydrogen gas, H_2?
(a) Same (b) five times
(c) 1/5 times (d) 1.5 times

x) The molar volume of CO_2 is maximum at:
(a) STP (b) 127°C and 1 atm
(c) 0°C and 2 atm (d) 273°C and 2 atm

Q.2. Fill in the blanks with suitable words given in the brackets:

i) The variable that stays constant when using the Charle's law is ________. (Pressure / Temperature)

ii) When the number of moles of a gas is ________, the volume will be doubled. (doubled/tripled)

iii) Partial pressure of oxygen in the ________ at one atm is 159 torr. (lungs/air)

iv) Gases have ________ shape and indefinite volume. (definite/indefinite)

v) Mathematical expression for Boyle's law is ________. ($PV=K$/$V \propto T$)

vi) Plasma is a ________ of electricity. (conductor/ non-conductor)

vii) The ________ has no relationship with plasma. (moon/aurora)

viii) The number of molecules of hydrogen and oxygen is ________ in one dm^3 of both gases at same temperature and pressure separately. (same/different)

ix) Charles' law established a relationship between volume and ________ in 1787. (pressure/temperature)

x) A gas can be liquefied by applying pressure when the temperature is ________ its critical temperature. (above/below)

Q.3. Label the following statements as True or False.

i) Gases contract on heating and expand on cooling.

ii) The collisions of ideal gas molecules are perfectly elastic.

iii) High temperature and low pressure make the gases non-ideal.

iv) The value of general gas constant at STP is 0.08206 atm $dm^3 mol^{-1} K^{-1}$.

v) The rate of diffusion varies directly to the square root of its molar mass.

vi) At constant temperature and pressure one mole each of N_2 and O_2 gases have same kinetic energy.

vii) Lighter gases can diffuse more quickly than heavier one.

viii) Ideal gas is the one which does not obey gas laws strictly at all conditions of temperature and pressure.

ix) Van der Waal's equation is used to calculate the pressure and volume of real gases.

x) The scale at which boiling point of water is taken is 212°F is known as Celsius scale.

Q.4: Define pressure and give the common units for pressure.

Q.5: What is barometer? How does it measure theatmospheric pressure?

Q.6: Name four elements and four compounds that exist as gases at room temperature.

Q.7: What are the basic assumptions of the kinetic–molecular theory of gases? Which assumption is incorrect at very high pressures? Which one is incorrect at low temperatures?

Q.8: Describe Boyle's law of gases? Give the mathematical expression and graphical representation of this law.

Q.9: Describe Charles' law of gases? Give the mathematical expression of this law.

Q.10: Explain absolute temperature scale on the basis of Charles Law.

Q.11: What is Avogadro's law of gases? Describe an experiment to demonstrate it.

Q.12: Derive ideal gas equation. Describe the applications of ideal gas equation.

Q.13: Calculate the numerical value of ideal gas constant R at STP for one mole of a gas:

(a) **When the pressure is in atmosphere and volume is in dm^3.**

(b) **When the pressure is in Pascal and volume is in m^3.**

Q.14: Define the molar volume, how will you calculate it?

Q.15: What is compressibility factor? The compressibility factor values can be less than or more than one for a real gas. Justify your answer.

Q.16: Describe the difference between an ideal gas and a real gas. Under what conditions does a real gas behave most like an ideal gas?

Q.17: (a) Why gases show non-ideal behaviour at low temperature and high pressure?

(b) Why helium and hydrogen are ideal whereas ammonia and sulphur dioxide are non-ideal at room temperature and one atmosphere pressure?

Q.18: Define Van der Waal's equation and derive it for real gases.

Q.19: What is meant by the partial pressure of a gas? Explain Dalton's law of partial pressure and describe its important applications.

Q.20: Define mole fraction. Does mole fraction have units?

Q.21: (a) What is the difference between effusion and diffusion?

(b) What are the factors on which the rate of effusion depends?

Q.22: Define and explain Graham's law of diffusion and effusion and describe experiment to demonstrate it.

Q.23: What is critical temperature? What is its importance for liquefaction of gases? Why a gas can't be liquefied above its critical temperature?

Q.24: What is meant by liquefaction of gases? Discuss the Lind's method of liquefaction of gases. What are the examples of liquefied gas in daily use?

Q.25: Define plasma and describe some of its properties. How can you differentiate between plasma and gas?

Q.26: Where in the universe are plasmas commonly found? Where can plasmas are found on Earth? What are the applications of plasma?

Q.27: Answer the following questions.

a) Why is mercury a more suitable substance to use in a barometer than water?

b) What is meant by standard temperature and pressure (STP)?

c) Why the Scientists prefer the Kelvin scale to the Celsius, and to the Fahrenheit, scales?

d) Why hydrogen and oxygen gases move with different velocities at the same temperature although their kinetic energies are same?

e) Explain why -273.15°C is the lowest possible temperature.

f) Explain why pressure increases as a gas is compressed into a smaller volume.

g) Why does a helium-filled balloon lose pressure faster than an air-filled balloon?

h) Explain why a helium filled weather rubber balloon expands as it rises in the air. Assume that the temperature does not change.

i) Two identical balloons have same volume of gases at STP. One balloon has nitrogen gas and the other has carbon dioxide gas. Which balloon has more molecules?

j) Why do deep-sea divers breathe a mixture of helium and oxygen?

k) When a gas is collected by the downward displacement of water, is a pure gas? If yes, then why, if not, then why not?

l) Explain why the second story of a double-story building is often warmer than a ground story during hot summer days?

(m) Which gas will a student smell first, the ammonia, NH_3, and hydrogen sulphide, H_2S? When both are released together across a class room?

(n) Why are the rates of diffusion and effusion of CO & N_2 the same?

(o) Which gas has stronger intermolecular forces, argon (T_c=-122.3 °C) or methane (T_c=-81.9°C) and why?

(p) The temperature of a highly compressed real gas that is allowed to expand into a vacuum usually falls. Why?

(q) We often experience pain in our ears because of rapid changes in altitude, why?

(r) Would it be easier to drink apple juice with a straw on top of Mt. Everest or at the base? Explain

Q.28: The pressure of helium gas in the neon sign is 7.6cmHg. What is the pressure of this gas in the neon sign in psi and mbar?

Q.29: At what pressure would neon gas occupy 2.85 dm^3 if it occupies 142.4 dm^3 at a pressure of 2.22 atm? (The temperature does not change).

Q.30: A Freon gas (used in the refrigeration systems) has a volume of 2.5 dm^3 at pressure of 76 mmHg. What is the pressure of this gas when it is expanded until its volume is 12.5 dm^3? Assume the temperature of gas does not change.

Q.31: A sample of gas at 10°C and 760mmHg contains a volume of 8.45 dm^3. What is the volume of this gas at 43°C and one atmosphere pressure?

Q.32: The volume of a bubble is 2.5mL. If its temperature increases from 30 to 70°C, what is its final volume?

Q.33: The volume of argon gas at 120°C is 380mL. The gas is heated to a volume of 560mL. What will be the new temperature of gas if the pressure is kept constant?

Q.34: What is the effect on the volume of a gas if you simultaneously:

a) Halve its pressure and double its Kelvin temperature?

b) Halve its pressure and halve its Kelvin temperature?

c) Double its pressure and double its Kelvin temperature?

Q.35: The weather balloon is inflated to a volume of 605 dm^3 at STP. The balloon is heated to 35°C. What would be its volume at 76cmHg?

Q.36: A 0.75 moles of hydrogen gas contains a volume of 16.8 dm^3. What is the volume of 1.25 moles of this gas at constant temperature and pressure?

Q.37: What is the volume occupied by 2.75 moles N_2 gas at STP?

Q.38: The helium filled balloons are used to record temperature, pressure, relative humidity, and wind velocity in the upper atmosphere. Supposed a balloon is launched at a room temperature (25°C) and 760mmHg. Its volume is $3.5m^3$ at ground level. What will be its volume at height of 30 miles where pressure is 40mmHg and temperature is –40°C.

Q.39: The volume of laughing gas (N_2O) at 480mmHg and 100°C is 800mL. The gas is cooled to a volume of 300mL at 760mmHg. What is its new temperature?

Q.40: The volume of nitrogen gas in an automobile air bag is 25 litre at –20°C and 1.50 atm. Calculate the number of moles of nitrogen gas in the air bag.

Q.41: A motor cycle tube contains 10.2 grams of gas and has a volume of 3×10^3mL Calculate the molar mass of the gas at 32°C and 506625 Nm^{-2}.

Q.42: What is the molar mass of a gas that has a density of 5.75 g/L at STP?

Q.43: What are the densities of N_2 and He at STP?

Q.44: What is the total pressure in a cylinder filled with air if the pressure of the oxygen is 250mmHg and the pressure of nitrogen is 510mmHg?

Q.45: A CNG cylinder contains a mixture of methane and ethane gases. The total pressure of the gases is 6.5 psi. What is the partial pressure of methane gas, if the partial pressure of ethane gas is 2.5 psi?

Q.46: Calculate the partial pressures of 2.5g Hydrogen, 5g Neon, and 15g Krypton at STP.

Q.47: Calculate the ratio of effusion rates of helium and nitrogen gases at same temperature and pressure.

Q.48: Oxygen gas effuses 1.173 times as swift as at the rate of an unknown gas through a balloon. Calculate the molar mass of unknown gas and identify it. This gas is produced during the combustion of hydrocarbons like methane.

Q.49: H_2 effused 2.82 times faster than an unknown gas. What was the molar mass of the unknown gas?