CHAPTER 6

STATE OF MATTER AND THE GAS LAWS

Matter is defined as anything that has mass and occupies space. Matter is everything on earth including the universe itself. They include.

1. Things too small to be seen with our naked eyes e.g. atoms, protons, electrons, ions, amoeba, etc
2. Things we cannot even see at all but can be detected by other means e.g. rays, gas, air, etc.

Experimental results have proved that matter contains positive and negative charges. Since matter appear to be uncharged, it then means that matter contain equal numbers of positive and negative charges.

6.1 STATE OF MATTER

The three physical states of matter are

1. SolidState: Here the particles are held in position by strong inter-particular (cohesion) force. Hence solid state matter

i. Have a fixed position
ii, Have a definite volume and shape
iii. Have only enough kinetic energy to vibrate
v. Not compressible.
vi.

2. LiquidState: Here the particles inter-particular (cohesion) forces are not as strong as in the solid state. Hence the liquid state matter

i. does not have a fixed position
ii. Have a definite volume but not a definite shape.
iii. Have enough kinetic energy to vibrate and translate.
iv. Not compressible.

3. Gaseous State: Here the particles are far apart and the forces of attraction between them are very weak. Therefore, the gaseous state matter:

I. Does not have a fixed position
II. Does not have a definite volume nor shape
III. Have enough kinetic energy to vibrate and translate
IV. Compressible.

It should be noted that when the heat content in matter is altered, matter changes from one state to another. Also, some matters changes from solid to gaseous state directly and vice versa without passing through the liquid state. Such matters are said to *sublime*.

6.2 THE GAS LAWS

The Gas Laws provides the description of the ideal gas behaviour based on experimental works by notable scientists like Boyle, Charles, Dalton and Graham.

6.2.1 THE KINETIC THEORY OF GASES

The Kinetic theory of gases states that:

i. Gaseous particles are very tiny.

ii. Gases have negligible force of attractions because their molecules are far apart.

iii. Gases also have negligible volume because their molecules arc so tiny, far apart, and with empty spaces between them, making their actual volume very little compared to their containing vessels.

iv. The gas molecules are always in a state of random motion hence the total kinetic energy of the entire gas molecule in a container at constant temperature and pressure is constant.

v. The pressure exerted by a gas depends on the contact collision of the gas molecules with themselves and with the walls of the containing vessels.

6.2.2 BOYLE'S LAW

Boyle's law was put forward by an Irish scientist Robert Boyle in the 1600s. It states that the volume of a given mass of gas is inversely proportional to its pressure provided temperature remains the same (constant). That is, as volume is increasing, the pressure will be decreasing at the rate in which the volume is increasing and vice versa. This means if the volume is doubled, the pressure will be halved. And the product of the two variables (i.e. volume and pressure) will always remain the same (i.e. constant).

The law can be expressed mathematically as V a $\frac{1}{P}$

$\therefore$ -V = $\frac{K}{P}$

K = VP

Where
V = Volume of the gas sample
P = Pressure exerted by tile gas sample.
K = Constant

This law can be shown to be true by confining a known volume of air in cylinder with a movable piston.As the volume of the gas is varied, the resultant change in pressure is read and recorded.

From the results, it will be observed that as volume is halved, pressure is doubled and vice versa. And the product of pressure and volume is always the same.

$V_1 P_1 = V_2 P_2 = V_3 P_3$

Table 6.1: **Experimental Data Illustrating Boyle's Law**

e	Volume	Pressure	Volume x Pressure
1	100.0cm^3	10.0 atm	1000 atm cm^{-3}
2	500.0cm^3	2.00 atm	1000 atm cm^{-3}
3	250.0cm^3	4.00 atm	1000 atm cm^{-3}
4	125.0cm^3	8.00 atm	1000 atm cm^{-3}
5	62.5cm^3	16.00 atm	1000 atm cm^{-3}

Figure 6.1: Graph illustrating Boyle's law (Gas pressure – Volume relationship). A plot of Volume versus Pressure

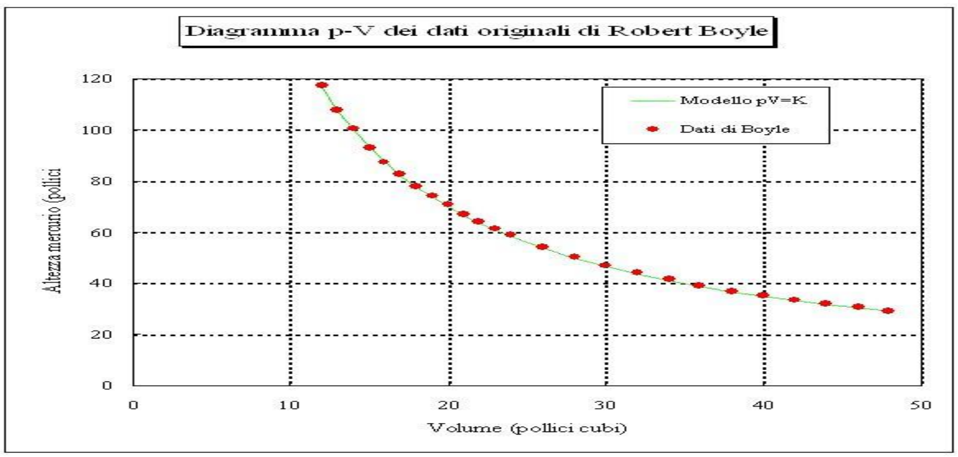
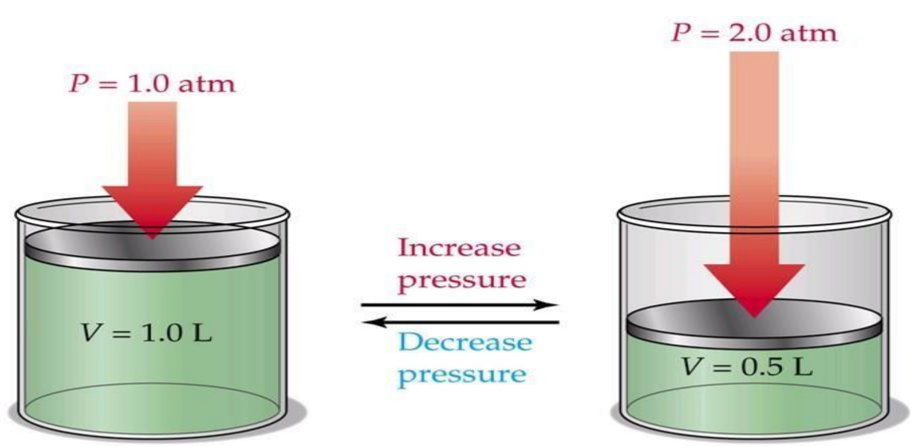


Figure 6.2: Boyle‟s law experiment (a) The volume at 1 atm pressure is 1.0L (1000cm^3) (b) When pressure is doubled to 2 atm, the volume will be halved (0.5L or 500cm^3). Tripling the pressure will reduce the volume to one-third of the original volume



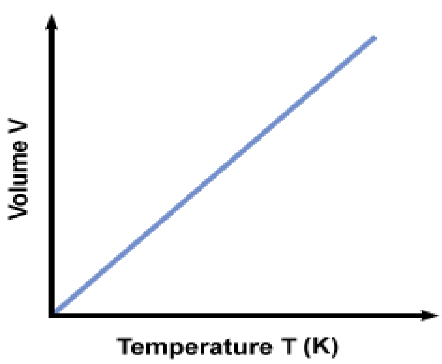
6.2.3 CHARLES LAW

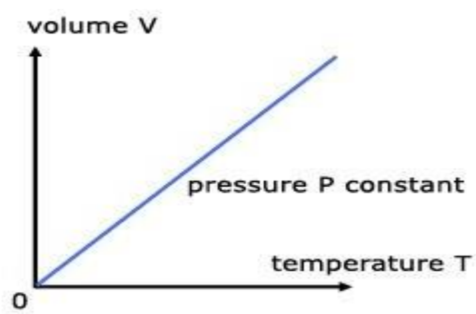
The relationship between gas volumes and temperature at constant pressure was put forward by a French scientist Jacques Charles in the l700s. The law states that: The volume of a given mass of gas is directly proportional to its absolute (Kelvin) temperature if the pressure remains the same (constant). That is as the volume is increasing, the temperature will equally increase and vice versa. And in every case, the value obtained by dividing the volume by the Kelvin temperature isalways the same.

The law can be expressed mathematically as $V \alpha T$
$\therefore V = KT$

Rearranging, we get
$\frac{V}{T} = K$ (At a constant pressure)

Where
V = Volume of the gas sample
T = Kelvin temperature of the gas sample
K = the constant





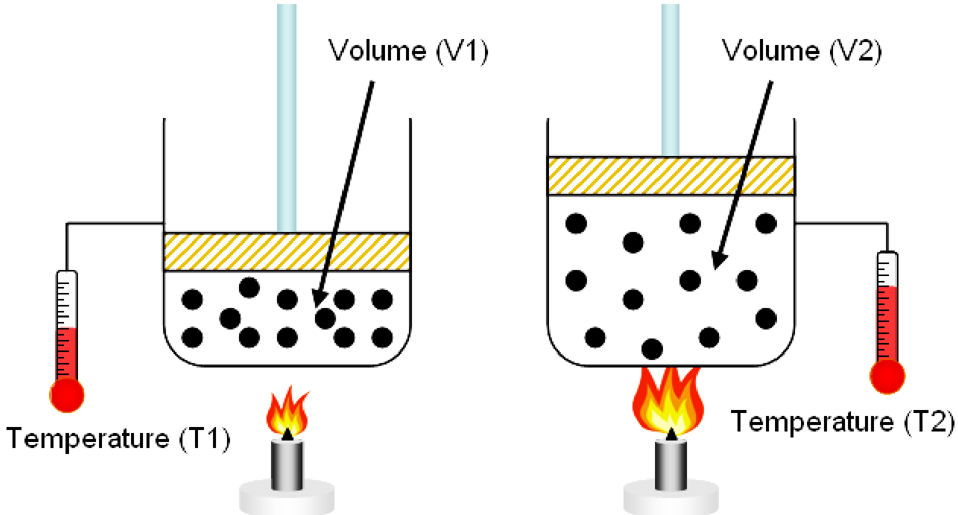
Charle‟s law
0 Kelvin (absolute zero)

The relationship between the gas volumes and temperature can be demonstrated by confining a gas in a cylinder with a weighted piston. Here, the pressure exerted by the gas is equal to the pressure exerted on the gas by the weighted piston which is not moving.

As the temperature is increased by heating, the gas will expand, forcing the weighted piston upward so as to give the gas an increased volume. Also, when the temperature of the cylinder is reduced by cooling, the gas will contract, forcing the weighted piston downward to assume a reduced gas volume

Figure 6.3

Charles experiment: At constant pressure, if temperature is doubled, the volume will be increased to two times the original volume.



From the results of the experiment, it is observed that at every change in temperature, the volume changed, in the same direction. That is, increased temperature gave increased volume and vice versa. And for every trial, the value obtained by dividing volume by Kelvin temperature was always, the same:

i.e. $\frac{V_1}{T_1} = \frac{V2}{T_2} = \frac{V3}{T_3}$

6.2.4 GENERAL GAS EQUATION

This is a mathematical statement, which combines Boyle's law and Charles law. This equation is used by chemist to investigate what the volume of gases will be when both their temperature and pressure are changed. It is used to solve problems in which all the three variables volume, pressure and temperature change.

From Boyle‟s law:

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

Combining both laws we get:

$V \alpha \frac{T}{P}$

$\therefore$ -V = K $\frac{T}{P}$ re-arranging we get:

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

$\therefore \frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2} = \frac{V_3P_3}{T_3}$

6.2.5 STANDARD TEMPERATURE AND PRESSURE (STP)

As the volume of a sample of gas depends on the temperature and pressure, therefore for comparing volumes of gases, a standard temperature and pressure are used. The standard temperature is 0°C (273K) while the standard pressure is 760 mmHg (101.3×10^3pa). The standard temperature and pressure can also be called STP

Example 6.1

A certain mass of gas occupies 100cm^3 at 255mmHg. What will be its volume at 390mmHg, assuming no change in temperature.

Solution

The question involves only the gas volume and pressure. Therefore, Boyle‟s law mathematical expression will be used in solving the problem.

$V_1P_1 = V_2P_2$

Initial gas volume V_1 = 100cm3

Initial gas pressure P_2 = 255mmHg

Final gas volume V_2 = unknown

Final gas pressure P_2 = 390mmHg

Make the unknown value subject of the formula

$$\frac{V_1P_1}{P_2} = V_2$$

$$\therefore V_2 = \frac{100\text{cm}^3 \text{ x } 255\text{mmHg}}{390\text{mmHg}}$$

$$V_2 = \frac{10200\text{cm}^3}{780} \qquad = 130.77\text{cm}^3$$

6.2.7 DALTON'S LAW OF PARTIAL PRESSURES

Dalton's law of partial pressures states that the total pressure of a mixture(non-reacting gases) of ideal gases is the sum of all the individual pressures of the gases only when they occupy the same volume:

$$P = P_A + P_B + P_C$$

Note that a mixture of gases contains two or more gases which do not: react chemically together. Also, the partial pressure is the pressure exerted by each of the gases in the mixture.

That is, the total pressure of a mixture of 1dm^3 of oxygen at 100mmHg and 1dm^3 of nitrogen at 200mmHg will be 300mmHg.

6.2.8 GRAHAM'S LAW OF DIFFUSION

Graham's law of diffusion state that at constant temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of its density. Molecular masses can also be used in terms of the density. Graham's law mathematical expression is:

$R_1 = \frac{K}{\sqrt{d_1}}$

$\frac{R_1}{R_2} = \sqrt{\frac{d_2}{d_1}}$ OR $\frac{R1}{R2} = \sqrt{\frac{m_2}{m_1}}$

R_1 and R_2: Initial and final rate of diffusion
d_1 and d_2: Initial and final densities of gases
m_1 and m_2: Initial and final molecular mass of gases

Example 6.2

What is the total pressure exerted by the mixture of l dm^3 of oxygen at 320mmHg. $1dm^3$ of nitrogen gas at 230mmHg

Solution

According to Dalton„s law of partial pressure, the total pressure will be the sum of all the individual partial pressures since they occupy the same volume.
$P = P_{oxygen} + P_{nitrogen}$
P = (320 + 230) mmHg = 530mmHg

Example 6.3

A volume of $3dm^3$ of nitrogen is mixed with a volume of $2dm^3$ of oxygen, with both at the same pressure of l01.3mmHg. What is the partial pressure of each gas in the mixture?

Solution

Gas A (nitrogen)

$V_1 = 3dm^3$ $P_1 = 101.3$
$V_2 = 5dm^3$ $P_2 = x$

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

$P_2 = \frac{101.3 \times 3}{5} = 60.78mmHg$

Gas A (Oxygem)

$V_1 = 2dm^3$ $P_1 = 101.3$

$V_2 = 5dm^3$ $\qquad$ $P_2 = x$

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

$P_2 = \frac{101.3 \text{ x}2}{5} = 41.32mmHg$

Example 6.4

A volume of nitrogen gas at 300 mmHg is 450cm^3. The gas is connected to oxygen at 400mmHgwith a volume of 550cm^3. What is the pressure of the mixture, assuming temperature is constant?

Gas A (nitrogen)

$V_1 = 450cm^3$ $\qquad$ $P_1 = 300mmHg$

$V_2 = 1000cm^3$ $\qquad$ $P_2 = x$

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

$P_2 = \frac{450cm3 \text{ X } 300mmHg}{1000} = 135mmHg$

Gas A (Oxygen)

$V_1 = 550cm^3$ $\qquad$ $P_1 = 400mmHg$

$V_2 = 1000cm^3$ $\qquad$ $P_2 = x$

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

$P_2 = \frac{550cm3 \text{ X } 400mmHg}{1000} = 220mmHg$

$P_T = P_{N2} + P_{O2}$

$P_T = 135mmHg + 220mmHg$

$= 355mmHg$

Examples 6.5

Calculate the rate of diffusion of methane (CH_4) and sulphur dioxide (SO_2).

Solution

The gases densities is not reflected in the question and in solving this problem the expression of Graham law that contains molecules masses showed be used because molecular‟ masses are easily determined from atomic masses.

$$\frac{R_1}{R_2} = \sqrt{\frac{m_2}{m_1}}$$

R_1 = Rate of diffusion of CH_4
R_2 = Rate of diffusion of SO_2
M_1 = Molecular mass of CH_4 = 16
M_2 = Molecular mass of SO_2 = 64

$$\frac{R_1}{R_2} = \sqrt{\frac{64}{16}}$$

$$\frac{R_1}{R_2} = \sqrt{4}$$

From the results, it shows that the rate of diffusion of methane CH_4 is twice that of sulphur dioxide.

PRACTICE QUESTIONS

1. What conditions are called the standard temperature and pressure (s.t.p) for measuring gases? (b) What is the molar volume of gases at s.t.p.?

2. The volume of a given mass of gas is $804cm^3$ at temperature of 127°C. Calculate the temperature of the gas when its volume is reduced to $603cm^3$ while the pressure remains constant.

3. A gas sample has a volume of $210dm^3$ at 37°C and 106.6mmHg. What would its final temperature be if the volume is reduced to $170dm^3$ and at a pressure of 76.4mmHg ;

4. A gas has a volume of $120dm^3$ at a pressure of 73.6mmHg and a temperature of 30°C. At what: pressure will the gas be when its volume is halved and at a constant temperature.

5. The maximum capacity of a container is $200cm^3$. If the container has a volume of $145cm^3$ at 20^oC, will an increase of temperature to 45°C increase the volume of the container to its maximum capacity assuming the pressure of the container is constant.

6. $600cm^3$ of nitrogen gas at 300mmHg is mixed with $400cm^3$ of oxygen at 500mmHg. Find the partial pressures of each gas and total pressure of the mixture when the temperature is kept constant·

7. (a) What gas law combines Boyle's Law and Charles Law?
   (b) Show its mathematical derivation