CHEMISTRY NOTES

FORM 3 SIMPLIFIED VERSION

QUICK REVISION NOTES

An Updated Well-Organized Detailed Revision Notes for the Current Form 3 Syllabus.

SERIES 1

THIS IS A FREE SAMPLE OF THE ORIGINAL NOTES

CONTACT US FOR FULL VERSION OF THE NOTES

Mr Isaboke 0746 222 000

MWALIMU CONSULTANCY

Printed in Kenya

00100

Nairobi

0746-222-000

mwalimuconsultancy@gmail.com

Mwalimu Consultancy Ltd.

GAS LAWS

Table of Contents

Specific Objectives

By the end of this topic, the learner should be able to:

- state the gas laws for an ideal gas
- verify experimentally the gas laws
- explain how the absolute zero temperature may be obtained from the pressure - temperature and volume - temperature graphs
- state the basic assumptions of the kinetic theory of gases
- explain the gas laws using the kinetic theory of gases
- solve numerical problems involving gas laws.

(15 Lessons)

Content

- Boyles law, Charles' law, pressure law, absolute zero
- Kelvin scale of temperature
  3. Gas laws and kinetic theory of gases(P= 1 ρ ‾ not required)
- Problems on gas laws [including $\frac{PV}{T}$ = constant]

GAS LAWS

An inflated balloon may burst when it gets warmer.The gas trapped inside the balloon is subject to changes in pressure, volume and temperature.

To explain why the balloon may burst, the relationships between these changes are investigated and they constitute what are termed as gas laws.

Boyle's Law

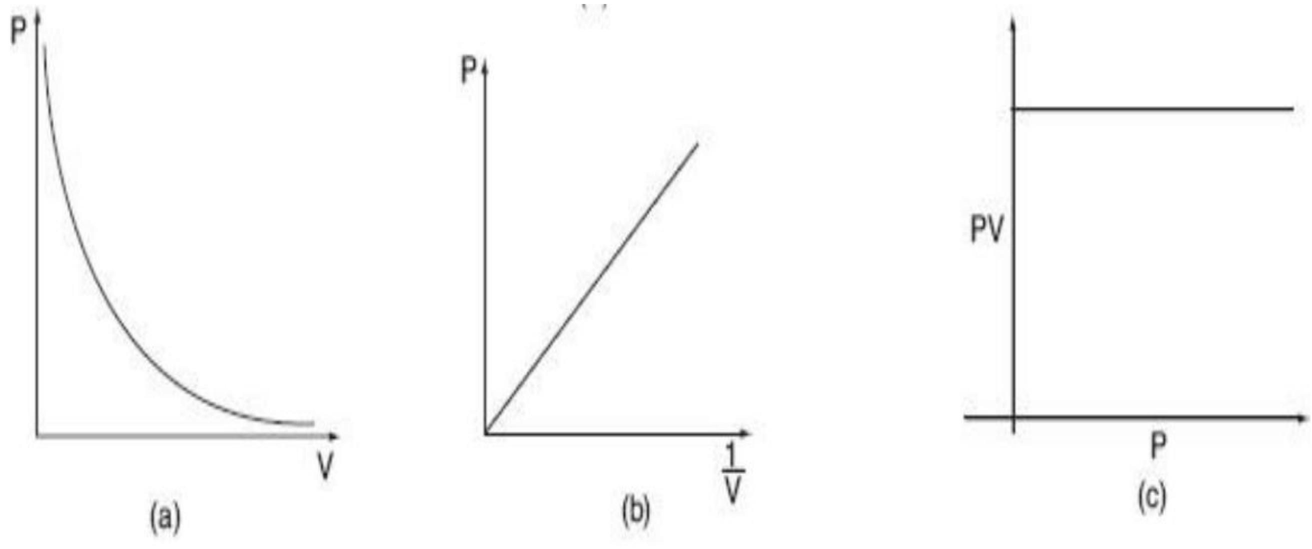
Boyle's law states that ***the pressure of a fixed mass of a gas is inversely proportional to its volume, provided the temperature is kept constant.***

Stated in symbols;

$$P \propto \frac{1}{V}, \text{ or } P = k \times \frac{1}{V}$$

So, PV = constant

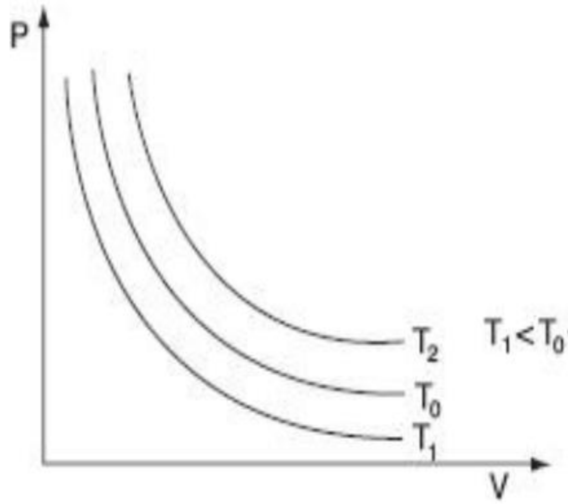
The sketches below show the relationship between pressure P and volume V of a fixed mass of gas.



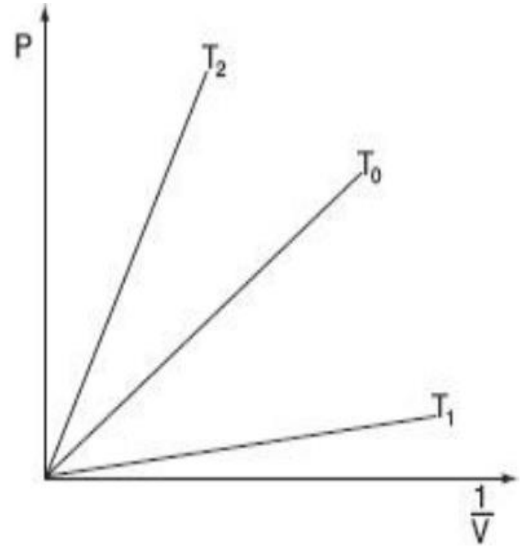
The graph of P against V is a smooth curve, as shown in (a), while that of P against $\frac{1}{V}$ is a straight line passing through the origin. That of PV against P is a straight line parallel to the x-axis. Since **PV = constant;**

$P_1V_1 = P_2V_2$ = constant, for any given mass of a gas.

At different temperatures, similar curves P against V are obtained as in. Each is called an **isothermal curve**.



(a) Pressure-Volume isothermal curve.



(b). P against 1/V for the isothermal curves.

When P is plotted against $\frac{1}{V}$ for each of the isothermals, the figure (b) above is obtained.

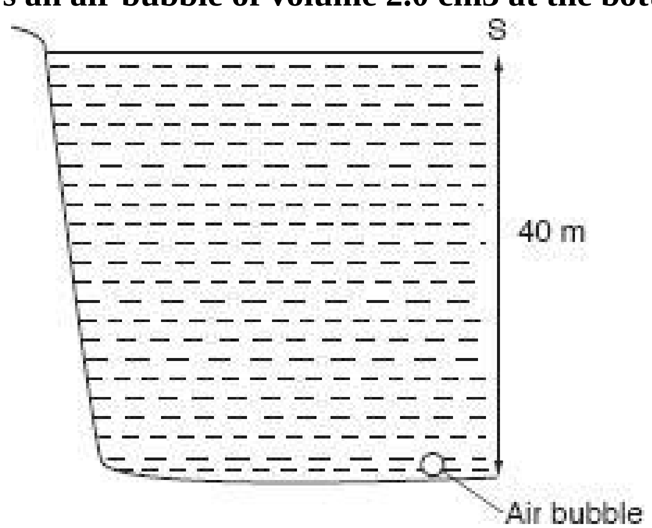
Using Kinetic Theory to Explain Boyle's Law

a) If the volume of a vessel containing a fixed mass of gas is **halved**, the number of molecules per unit volume will be **doubled**. The number of collisions per unit time, and therefore the rate of change of momentum, will also be **doubled**. Consequently, halving the volume of the gas doubles the pressure, which is the import of the Boyle's law.

b) In a demonstration using a syringe, as volume of gas is reduced, there is increase in number of molecular collisions resulting to increase in pressure.

Example 1

The diagram below shows an air bubble of volume 2.0 cm3 at the bottom of a lake 40 m deep.



Determine the volume just below the surface S if the atmospheric pressure is equivalent to a height of 10 m of water.

Solution

10 m height = 1 atm.
40 m height = 4 atm.
Pressure P1 at the bottom = (1 + 4)

- 5 atm

Pressure P2 at surface = 1 atm Volume V1 at bottom = 2 cm3 By Boyle's law, P1V1 = P2V2
5×2=1×V2 V2=10
Volume just below surface is 10 cm3.

Example 2

The volume V of a gas at pressure P is reduced to $\frac{3}{8}$V without change of temperature. Determine the new pressure of the gas.

Solution

PV = constant
P1V1 = P2V2

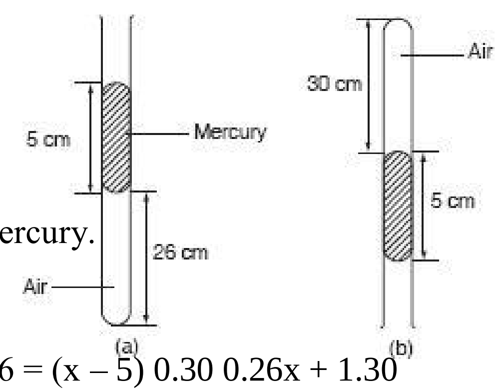
$$P_2 = \frac{8}{3}P_1$$

P1V1=P2× $\frac{3}{8}$V

The new pressure of the gas is $\frac{3}{8}$P.

Example 3

A column of air 26 cm long is trapped by mercury thread 5 cm long as shown in (a). When the tube is inverted as in (b), the air column becomes 30 cm long. Determine the value of atmospheric pressure.



Solution

In (a), gas pressure = atm pressure+ hρg.
In (b), gas pressure = atm. pressure – hρg, where ρ is the density of mercury.
From Boyle's law;
P1V1 = P2V2
Let the atmospheric pressure be height 'x' of mercury. So, (x + 5) 0.26 = (x – 5) 0.30 0.26x + 1.30 = 0.3x – 1.5

$$\therefore x = \frac{2.8}{0.04}$$

2.8 = 0.04x
= 70 cm

Review Exercise 1

1. In an experiment to verity Boyle's law, two quantities were advised to be kept constant

(a). State the quantities.

(b). the results of experiment to verify Boyle's law were recorded in the table below.

Pressure(atmospheres)	1.0	1.2	1.4	1.6	1.8
Volume (litres)	0.62	0.521	0.450	0.391	0.351

Plot a suitable graph to verify the law.

(c). Determine the volume of the gas when the pressure is two atmospheres.

2. (a) State Boyle's law

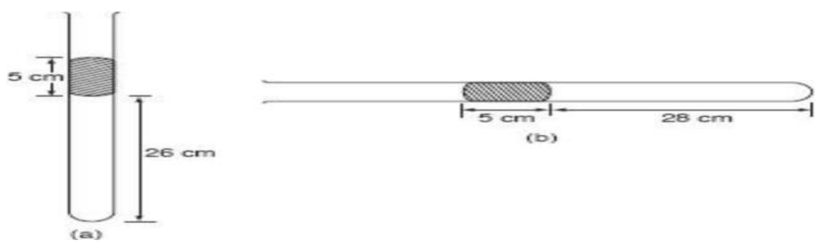
(b) The volume of a bubble at the base of a container of water is 3cm3. The depth of water is 30cm.The bubble rises up the column until the surface ;

(i) Explain what happens to the bubble as it rises up the water column

(ii) Determine the volume of the bubble at a point 5cm below the water surface

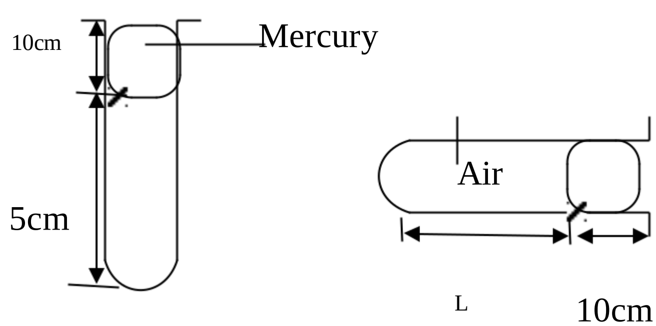
(c) A faulty thermometer records 11oC instead of 0oC and 98oC instead of 100oC. Determine the reading on the thermometer when dipped in liquid at a temperature of 56oC

a) A column of air 26 cm long is trapped by mercury thread 5 cm long as shown in diagram (a) below. When the tube is laid horizontally as in (b), the air column is now 28 cm. Find the atmospheric ppressure..



1. (a) State Boyle's law

(b) A column of air 5cm is trapped by mercury thread of 10cm as shown in the figure below. If the tube is laid horizontally as shown in **(b)**, calculate the new length of trapped air (atmospheric pressure =75.0cmHg and density of mercury = 13600kgm-3)



Charles' Law

Charles law relates the volume of a gas with its absolute (or Kelvin) temperature.

Charles' law states that ***the volume of a fixed mass of gas is directly proportional to its absolute temperature if the pressure is kept constant***.

In symbols, Charles' law can be stated as follows; V α T or V = kT, where k is a constant.

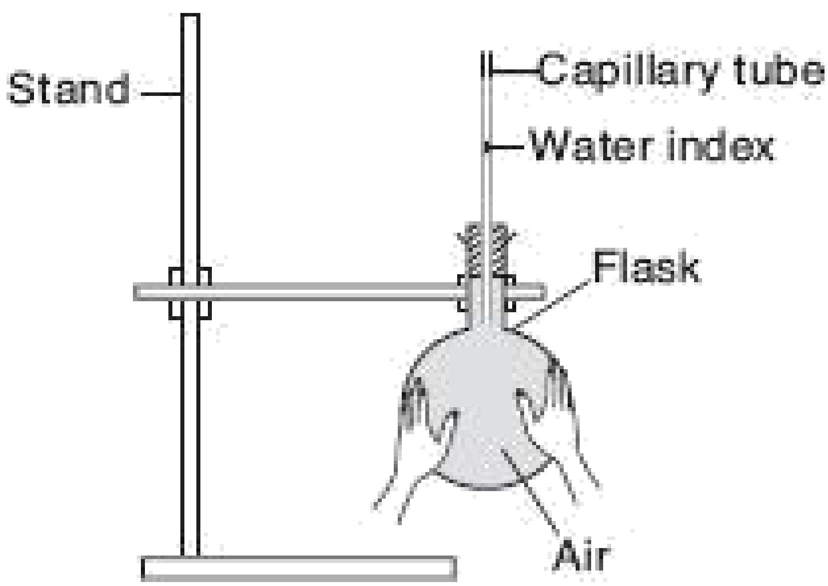
Hence, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ = constant

This formula is only applicable when T is expressed in Kelvins.

Experiment to verify Charles law

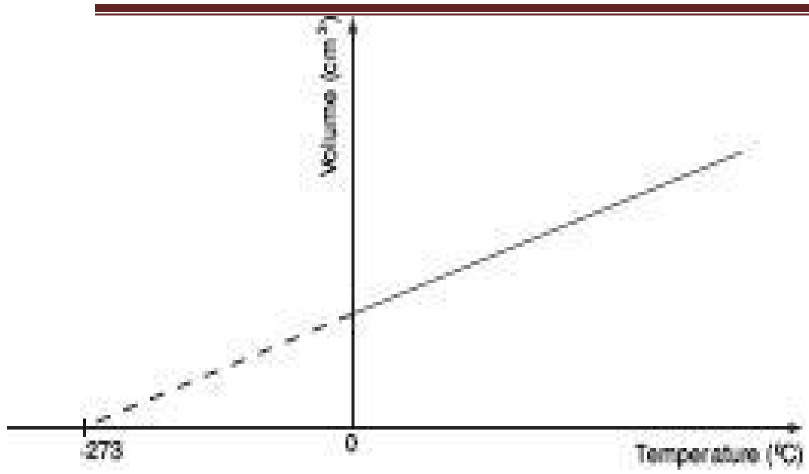
The set-up below can be used to show how temperature affects volume of a given mass of a gas at constant pressure.

The flask is grasped firmly and the water index observed.



The water index **rises higher when the flask is held and falls when the hands are withdrawn, showing that the volume of gas increases when its temperature is raised.**

As the temperature rises, the height h (volume) also increases. A plot of volume against temperature is represented alongside.

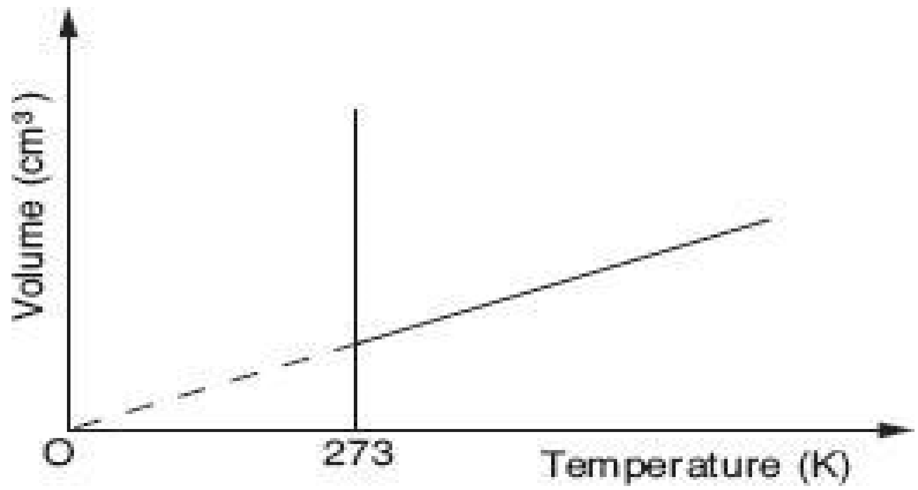


The graph is a straight line, indicating proportional changes in volume and temperature. However, it does not pass through the origin.

If the graph is extrapolated, it cuts the temperature axis at about –273°C. At this temperature, the volume of the gas is assumed to be zero.

This temperature, –273°C, at which the volume, pressure of the gas and kinetic energy of the particles are assumed to be zero is ideally the lowest temperature a gas can attain and therefore called **absolute zero.**

A plot of volume against absolute temperature gives a straight line through the origin, as shown below.



It follows that the volume of the gas is directly proportional to its absolute (or Kelvin) temperature
It is impossible to get to
NOTE: Absolute zero for gases because they condense (liquify) at fairly higher temperatures.

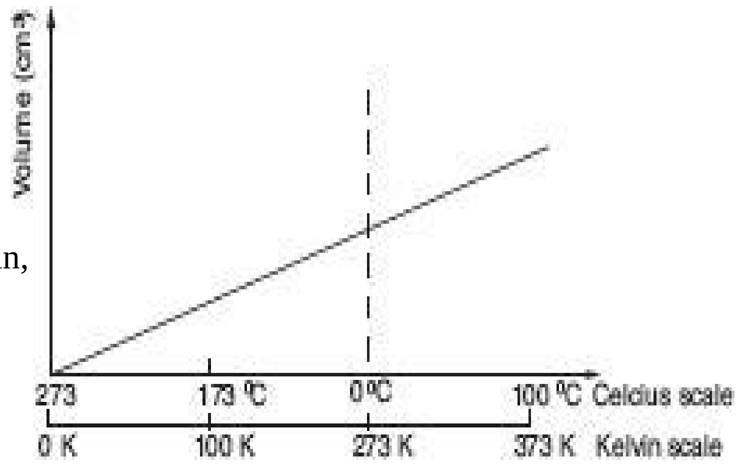
Using Kinetic Theory to explain Charles' Law

- When a gas is heated, the **kinetic energy and the velocity of the molecules increases.** As the temperature rises, the molecules **move faster.**
- If the volume of the container were constant, the pressure resulting from the collisions of the molecules with the walls would **increase due to greater rate of change of momentum per unit time.**
- But since Charles' law requires that the pressure be constant, then **the volume must increase** accordingly so that although the molecules are moving faster, the **number** of collisions at the walls of the container per unit time is **reduced**, since the distance between the walls is increased by increasing the volume.

Relation Between Celsius and Absolute scale

The zero Kelvin (0 K) corresponds to –273°C

while 0°C corresponds to 273 K.



It follows that to change from Celsius to Kelvin, we add 273 to the Celsius temperature, i.e.;
θ°C=T
=(θ+273)K

Example 1

The temperature of a gas is –42°C. What is this temperature on the Kelvin scale?

Solution

Temperature T = (– 42 + 273) K
=231K

Example 2

0.02 m3 of a gas at 27°C is heated at constant pressure until the volume is 0.03 m3.
Calculate the final temperature of the gas in °C.

Solution

$\frac{V}{T}$ = constant (pressure being constant)

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

$$\frac{0.02}{300} = \frac{0.03}{T_2}$$

$$T_2 = 300 \times \frac{0.03}{0.02}$$ change to celcius scale θ

- 450 K
- (450 – 273) °C
- 177°C

Example 3

A mass of air of volume is 750 cm3 is heated at constant pressure from 10°C to 100°C. Determine the final volume of the air.

Solution

Review Exercise 2

- (a) State: (i) Boyle's Law
  - Charles' Law.

(b) A form three student carried out an experiment on one of the gas law. She obtained the following results.

Temperature (^{0}c)	10	35	60	80	90	110
Volume V(cm^3)	5	5.8	6.4	7.0	7.2	7.8

i) Plot graph of volume V gainst temperature.

(ii) From the graph, determine the volume of the gas at 0oc.

(iii) Determine the slope of the graph.

(iv) The equation of the line obtained is of the form V = kT + c. What is the value of k and c?

i. A mass of gas occupies a volume of 150cm^3 at a temperature of -73°C and a pressure of 1 atmosphere. Determine the 1.5 atmospheres and the temperature 227 °C

PRESSURE LAW

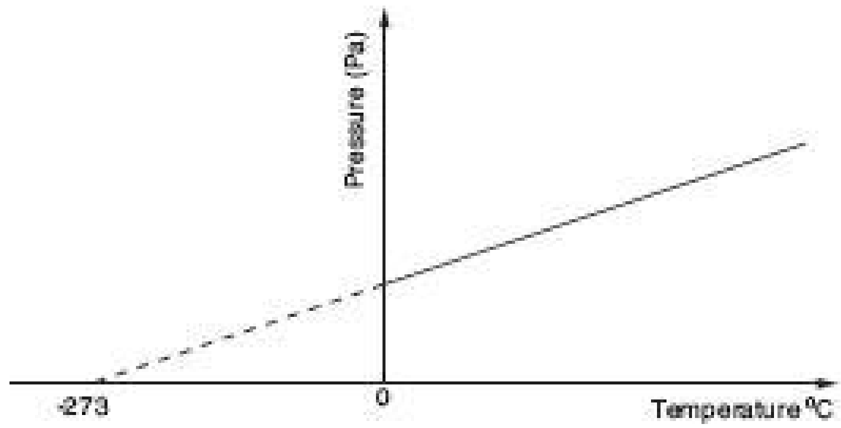
This law relates pressure of fixed mass of a gas to its absolute temperature at constant volume. Pressure law states **that the pressure of a fixed mass of gas is directly proportional to its absolute temperature, provided the volume is kept constant.**

In symbols;

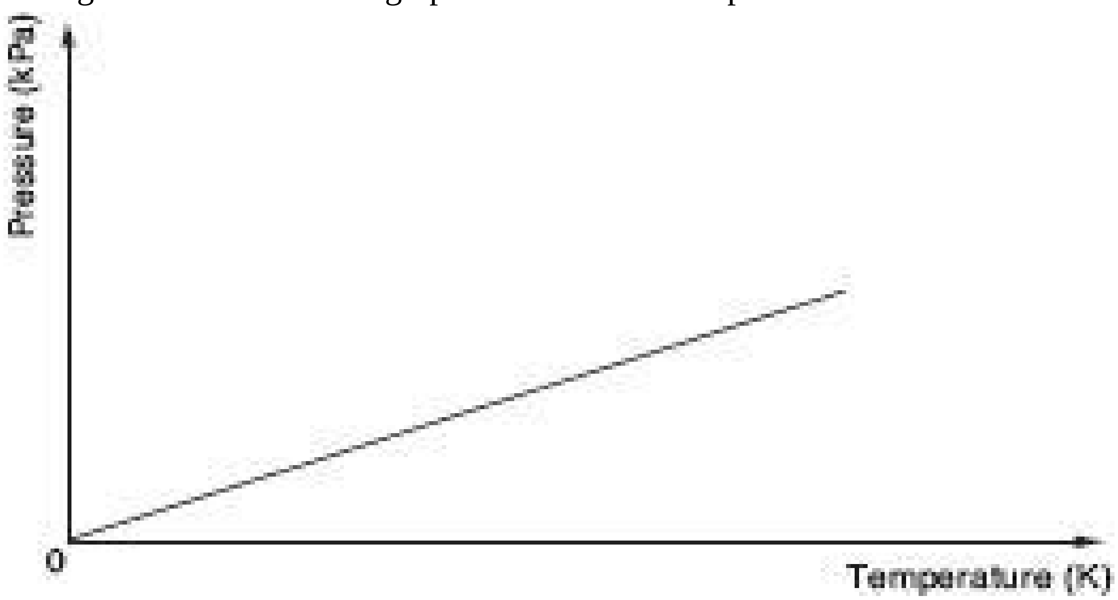
P α T (V constant)

Or P = kT, where k is constant

So, $\frac{P_1}{T_1} = \frac{P_2}{T_2}$



A plot of pressure against temperature gives the graph shown alongside.
When the graph is extrapolated, it cuts the temperature axis at –273°C, the absolute zero.
The figure alongside shows the same graph of an absolute temperature scale.



On the absolute scale, the pressure of a gas is directly proportional to its absolute temperature.

Using Kinetic Theory to Explain Pressure Law

- In gases, pressure is as a result of **bombardment** of the walls of the container by the gas molecules. When the molecules of the gas bombard and rebound from the walls of the container, a change of momentum takes place.

a) The number of bombardments per unit time constitutes a rate of change of momentum, which according to Newton's second law of motion, constitutes a force. This force per unit area emerges as the pressure of the gas.

b) When a gas is heated, its molecules gain kinetic energy and move about faster. If the volume of the container is constant, the molecules will bombard the walls many more times per unit time, and with greater momenta. The total rate of change of momentum will therefore increase. The resulting force per unit area, which is the pressure, will increase.

Example 1

A cylinder contains oxygen at 0°C, and 1 atmosphere pressure. What will be the pressure in the cylinder if the temperature rises to 100°C?

Solution

$$\frac{P}{T} = \text{constant}$$

$$\frac{1}{273} = \frac{P_2}{273}$$

$$P_2 = \frac{373}{273}$$

$$= 1.37 \text{ atmosphere}$$

Example 2

At 20°C, the pressure of a gas is 50 cm of mercury. At what temperature would the pressure of the gas fall to 10 cm of mercury?

Solution

$$\frac{P}{T} = \text{constant}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{50}{293} = \frac{10}{T_2}$$

$$T_2 = \frac{2\,930}{50}$$

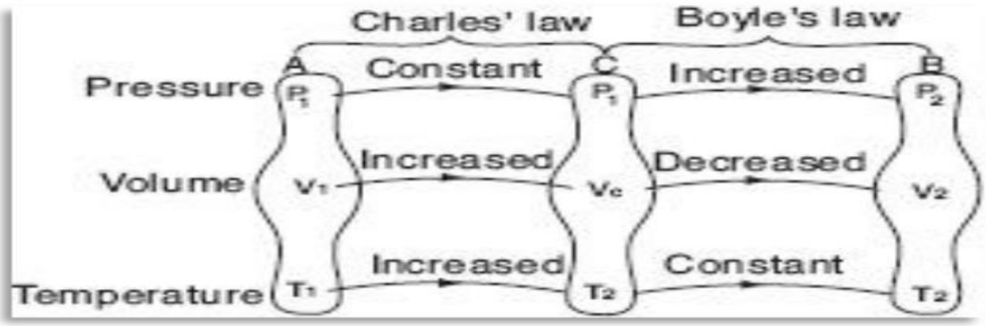
$$= 58.6 \text{ K} \quad (\text{or } -214.4°\text{C})$$

Equation of State

A general gas law relating the changes in pressure, volume and the absolute temperature can be derived from the three gas laws.

Consider a fixed mass of gas which is being changed from state A to state B through an intermediate state C, as shown below .

From A to C, the gas is heated at constant pressure P1. By Charles' law;



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

$$\text{Volume } V_c \text{ in state C, } V_c = \frac{V_1T_2}{T_1}$$

By Boyle's law, P1Vc = P2V2

$$V_2 = \frac{P_1V_c}{P_2}$$

$$\text{But } V_c = \frac{V_1T_2}{T_1}$$

$$\therefore\ V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$\text{Re-arranging, } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

In general, $\frac{PV}{T} = k$, where k is a constant.

This is known as the **equation of state**, in which k depends on the type and quantity of the gas.

The equation changes to $\frac{PV}{T} = R$ when the amount of gas is 1 mole. Constant R is same for all gases, and is called the **universal gas constant**.

Example

A mass of 1 200 cm^3 of oxygen at 27°C and a pressure 1.2 atmosphere is compressed until its volume is 600 cm^3 and its pressure is 3.0 atmosphere. Calculate the temperature of the gas after compression in °C.

Solution

V1 = 1 200 cm3

V2 = 600 cm3

T1=27+273

Q. 300 K T2=?

P1 = 1.2 atmosphere P2 = 3.0 atmosphere

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

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$= \frac{3 \times 600 \times 300}{1.2 \times 1\,200}$$

$$= 375\ k$$

$$= 102°C$$

Assumptions of Gas laws

- When explaining the gas laws using the kinetic theory, **both the size of molecules and the intermolecular forces are assumed to be negligible.**

- Real gases have molecules with definite volumes and therefore the idea of zero volume or zero pressure is not real. **Real gases get liquified before zero volume is reached.**

a) This departure from the gas laws is so particularly true at low temperatures and high pressures. A gas that would obey the gas laws completely is called **ideal** or **perfect** gas.

Recently…..

1. Figure 14 shows a graph of vapour pressure against the temperature of water vapour, in a laboratory where a mercury barometer indicates a height of 61.8 cm.

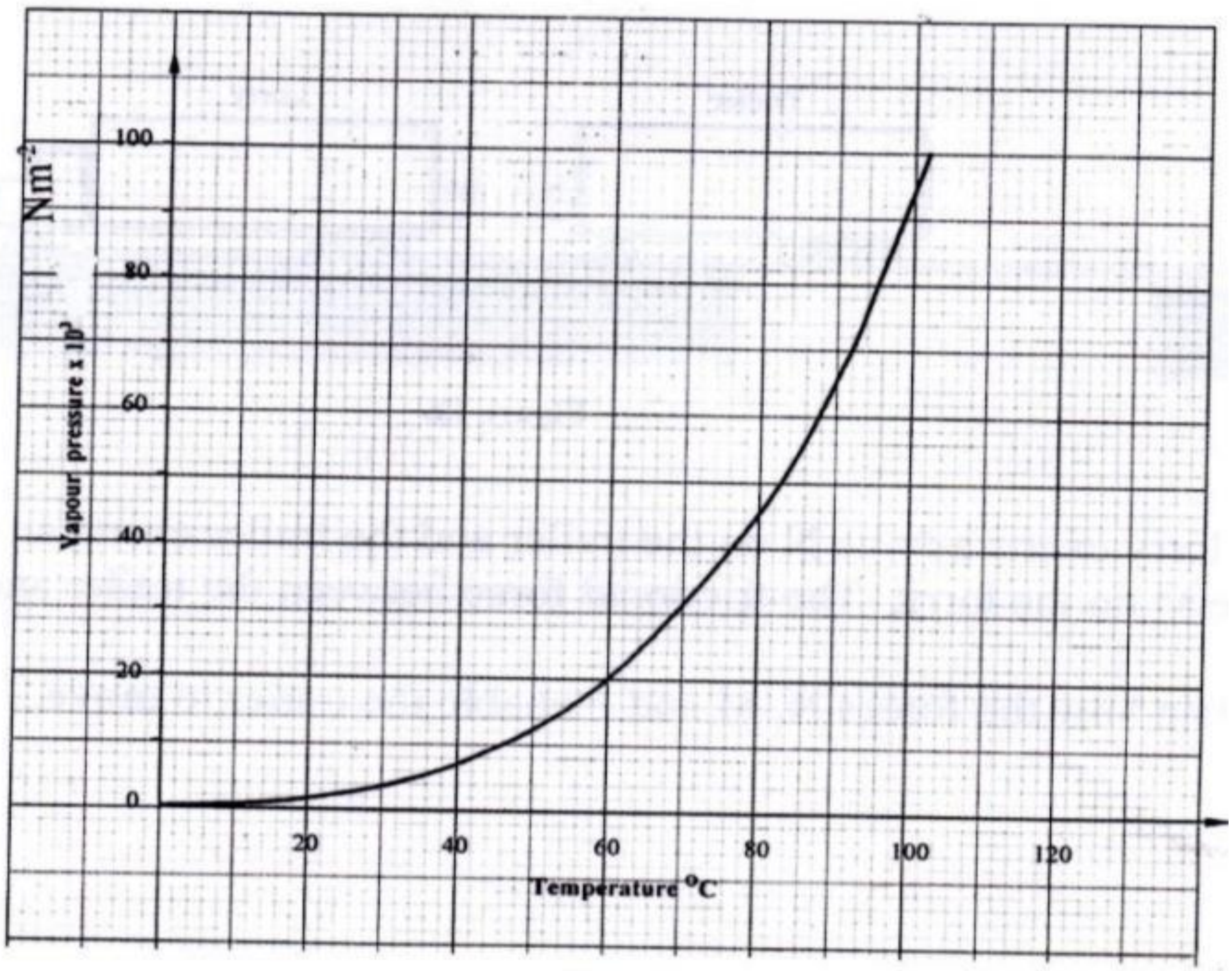


Figure 14

(i) Determine the atmospheric pressure in the laboratory in Nm -2

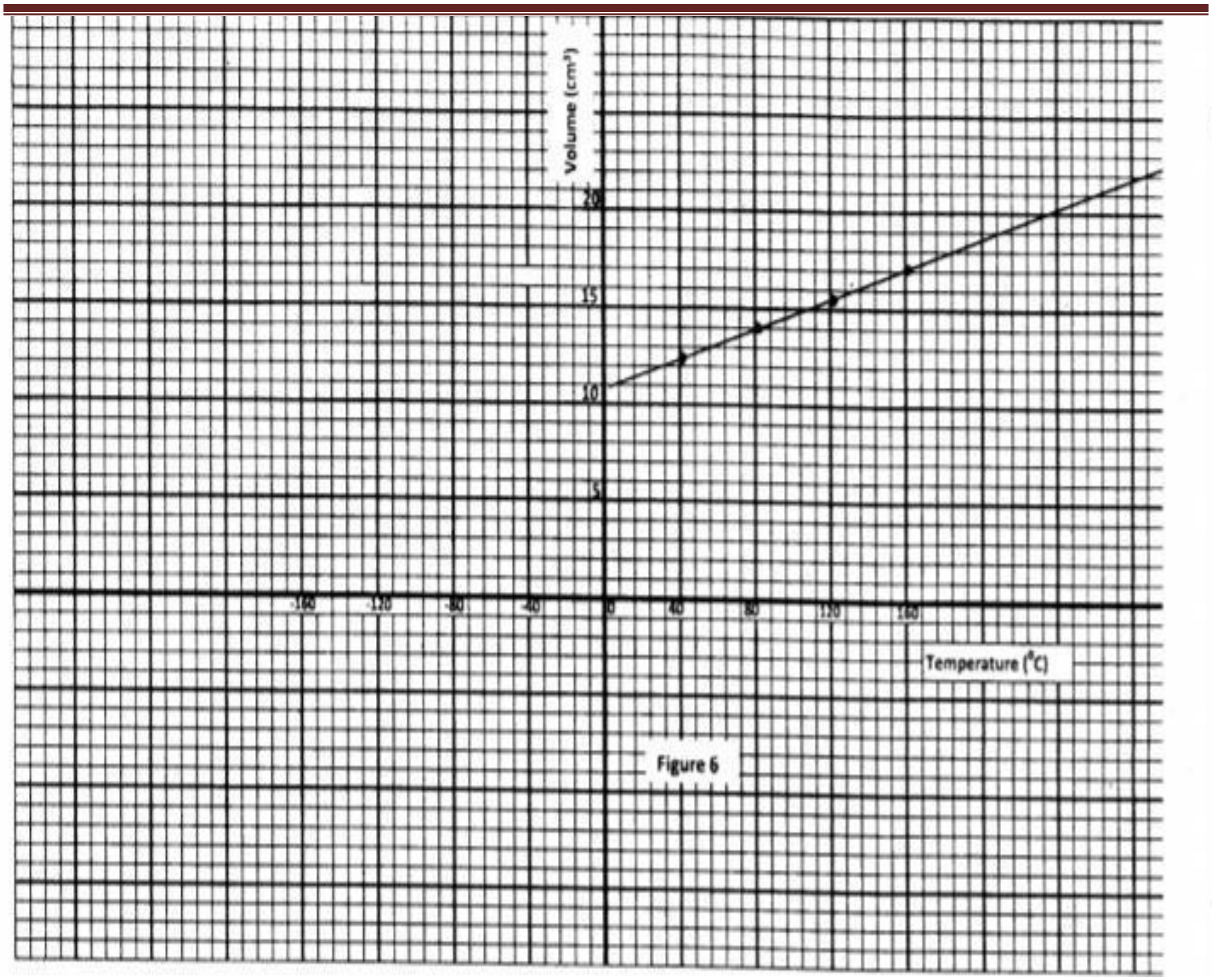
(Take g - 10m/S2and density of mercury =13600 kg/m3). (3 mks)

ii) Use the graph to determine the boiling point of water in the laboratory. (1mk)

i) A balloon is filled with hydrogen gas and then released into the air. It is observed that as it rises higher into the air it expands. Explain why it expands.

ii) In verifying the pressure law of gases, the temperature and pressure of a gas are varied at constant volume. State the condition necessary for the law to hold. (1 mk)

iii) Figure 6 shows a graph of volume against temperature for a given mass of gas.



Use the graph to determine the absolute zero temperature in °C. **(2 mks)**

Figure 7 shows a horizontal tube containing air trapped by a mercury thread of length 24cm.

- The length of the enclosed air column is 15cm. The atmospheric pressure is 76cmHg.

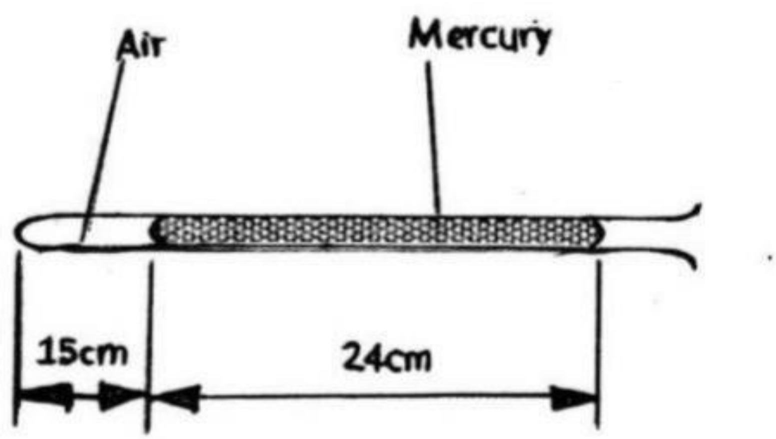


Figure 7

(a) State the pressure of the enclosed air. **(1 mk)**

- The tube is now held in a vertical position with the open end facing upwards as shown in Figure 8.

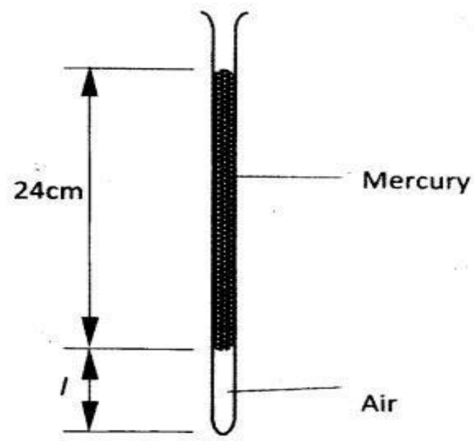


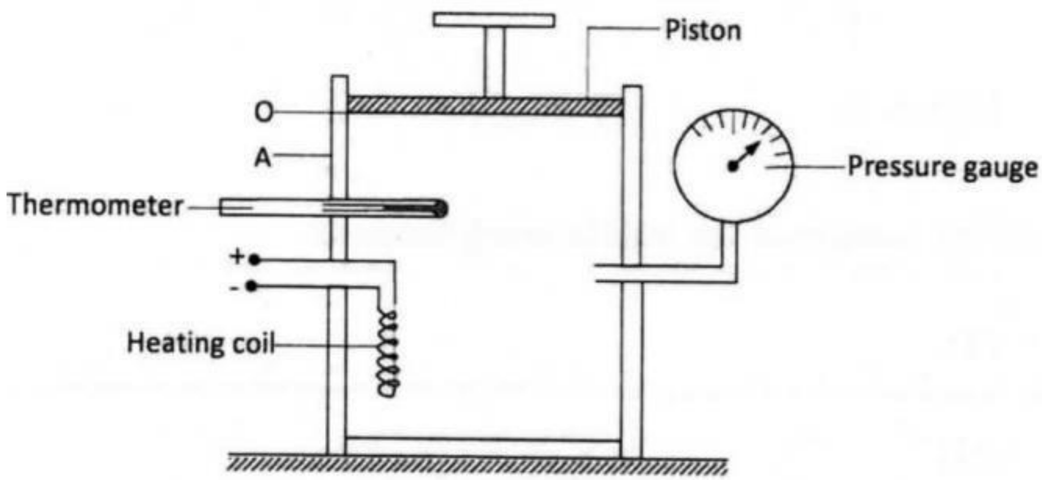
Figure 8

Determine:

-The pressure of the enclosed air.

-The length (l) of the enclosed air column

- **Figure 11** shows an insulated cylinder fitted with a pressure gauge, a heating coil and a frictionless piston of cross – sectional area 100cm2.



a) While the piston is at position O, the pressure of the enclosed gas is 10 Ncm2 at a temperature of 270C. When a 10 kg mass is placed on the piston, it comes to rest at position A without change in the temperature of the gas.

(i) Determine the new reading on the pressure gauge (4 mks)

- State with a reason how the value obtained in (i) compares with the initial pressure. (2 mks)

(b) The gas is now heated by the heating coil so that the piston moves back to the original position O.

i. State the reading on the pressure gauge. (1 mk)

ii. Determine the temperature of the gas in 0C. (4 mks)

(*Take g = 10 Nkg-1*)

a) A horizontal capillary tube of uniform bore sealed at one end contains dry air trapped by a drop of mercury. The length of the air column is 142 mm at 170C. Determine the length of the air column at 250C. (3 mks)

- The pressure of the air inside a car tyre increases if the car stands out in the sun for some time on a hot day. Explain the pressure increase in terms of the kinetic theory of gases. (3 mks)

- a) Figure 11 shows a graph of pressure (p) against volume (v) for a fixed mass of a gas at constant temperature

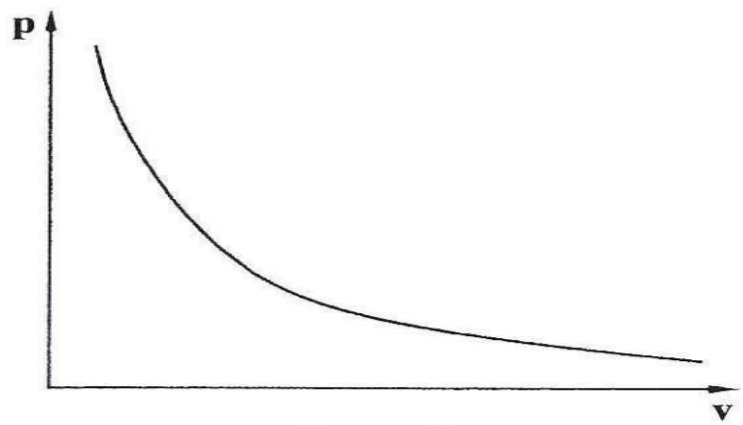


Figure 11

In the space provided, sketch the corresponding graph of P against 1/V(1 mk)
b)Explain the pressure law using the kinetic theory of gases(3 mks)

1. 20cm3 of a gas exerts a pressure of 760mmHg at 250C. Determine the temperature of the gas when the pressure increases to 900mmHg and the volume reduces to 15cm3 (4 mks)

10. (a) State the meaning of the term ideal gas. (1mk)

- The pressure acting on a gas in a cylinder was changed steadily while the temperature of the gas was maintained constant. The value of volume V of the gas was measured for various values of pressure.

The graph in **Figure 11** shows the relation between tin pressure P, and the reciprocal of volume, 1/V

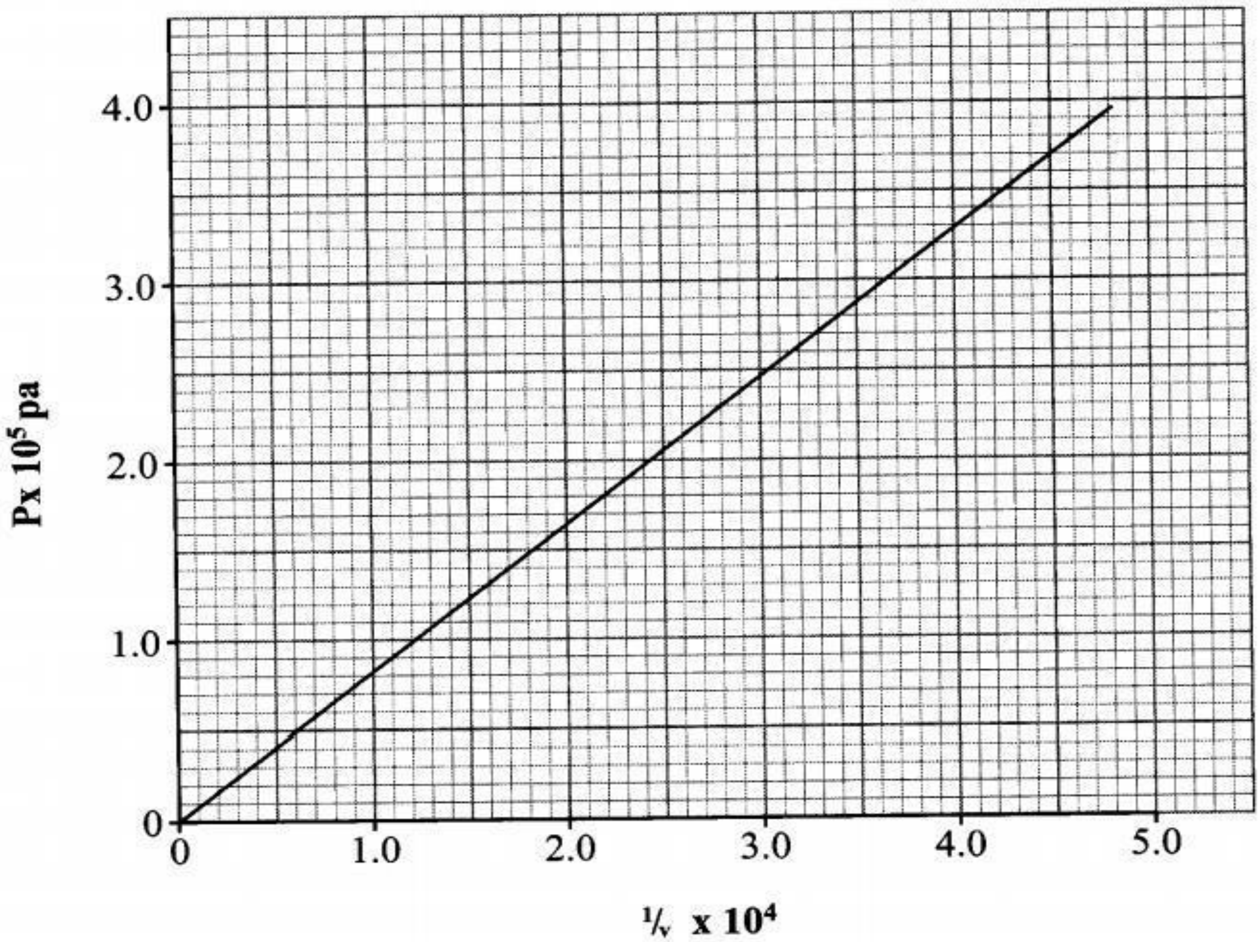


Figure 11

i) Suggest how the temperature of the gas could be kept constant. (2 mks)

ii) Given that the relation between the pressure ***P*** and the volume, V, of the gas is given by PV = *K*, where K is a constant, use the graph to determine the value of *K* (4 mks)

iii) What physical quantity does *K* represent? (1 mk)

a) State one precaution you would take when performing such an experiment.

(c) A gas occupies a volume of 4000 liters at a temperature of 37°C and normal atmospheric pressure. Determine the new volume of the gas if it is heated at constant pressure to a temperature of 67 °C (Normal atmospheric pressure, P = -1.01 X 105 Pa).

(4 mks)

- State one assumption for the experiments carried out to verity the gas laws. (1 mark)

- Figure 6 shows the relationship between volume and pressure for a certain gas.

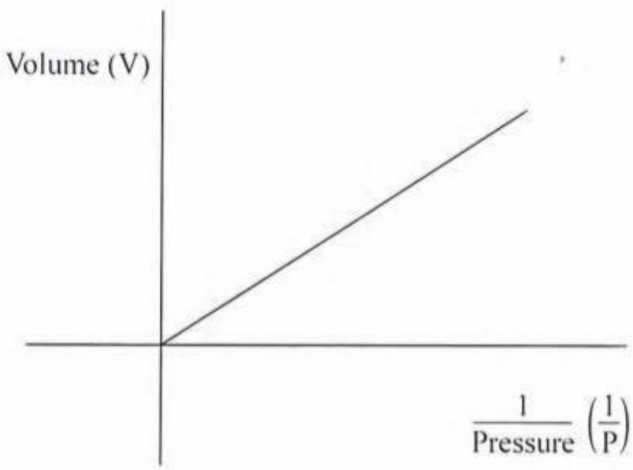


Figure 6

13. . (a) State two quantities that must be kept constant in order to verify Boyle's law. (2 marks)

i) An air bubble at the bottom of a beaker full of water becomes larger as it rises to the surface. State the reason why;

ii) the bubble rises to the surface,(1 marks)

iii) it becomes larger as it rises.(1 marks)

iv) State two assumptions made in explaining the gas laws using the kinetic theory of gases. (2 marks)

v) Figure 11 shows an incomplete experimental set up that was prepared by a student to verily one of the gas laws.

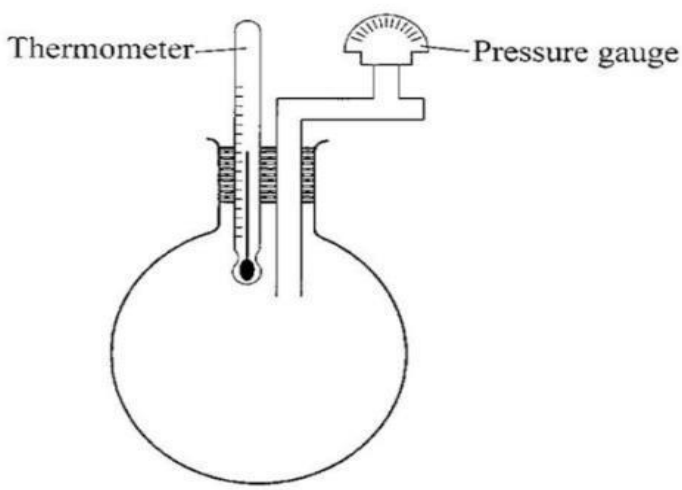


Figure 11

- State with a reason which one of the laws may be verified using the set up. (2 marks)
- State what the student left out in the diagram of the set up. (1 mark)

o The volume of a fixed mass of a gas reduced from 500 cm3 to 300 cm3 at constant pressure. The initial temperature was 90K. Determine the final temperature. (3 marks)

1. a) Figure 9 shows a graph of pressure against temperature for a fixed mass of gas at constant volume

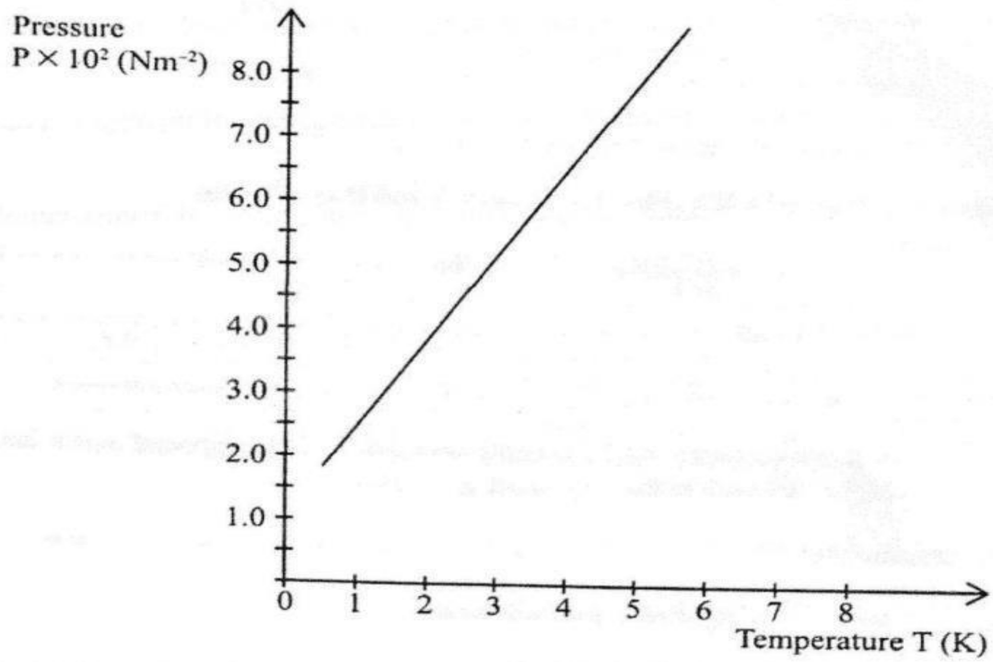


Figure 9

i) From the graph , determine that values of n and c given that P = nT + c where n and c are constant

ii) Explain why it is not possible to obtain zero pressure of a gas in real life situation **(2marks)**

iii) A fixed mass of a gas occupies 1.5 x10-3 m3 at a pressure of 760mmHg and temperature of 273 K. Determine the volume the gas will occupy at a temperature of 290K and a pressure of 720mmHg **(3marks)**

(d) state any three assumptions made in kinetic theory of gases **(3marks)**

(A)GAS LAWS

1. Matter is made up of small particle in accordance to Kinetic Theory of matter:
Naturally, there are basically **three** states of matter: **Solid, Liquid** and **gas:**
(i)A solid is made up of particles which are very closely packed with a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

(ii) A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

(iii)A gas is made up of particles free from each other. It thus has no definite /fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

2.Gases are affected by **physical conditions**. There are **two** physical conditions:
(i)**Temperature**
(ii)**Pressure**

3. The SI unit of temperature is **Kelvin**(**K**).
Degrees Celsius/Centigrade(**^oC**) are also used.
The two units can be interconverted from the relationship:

$^oC + 273 = K$

$K - 273 = {}^oC$

Practice examples
1. Convert the following into Kelvin.
(i) O **^oC**
^oC + 273 = K substituting : O ^oC + 273 = **273 K**

(ii) -273 **^oC**
^oC + 273 = K substituting : -273^oC + 273 = **0 K**

(iii) 25 **^oC**
^oC + 273 = K substituting : 25 ^oC + 273 = **298 K**
(iv) 100 **^oC**
^oC + 273 = K substituting : 100 ^oC + 273 = **373 K**

2. Convert the following into degrees Celsius/Centigrade(^oC).
(i) 10 K
K -273 = ^oC substituting: 10 – 273 = **-263 ^oC**

(ii) (i) 1 K

K -273 = ºC substituting: 1 – 273 = **-272 ºC**

(iii) 110 K

K -273 = ºC substituting: 110 – 273 = -163 ºC

(iv) -24 K

K -273 = ºC substituting: -24 – 273 = -297 ºC

The **standard** temperature is **273K** = **0 ºC**.
The **room** temperature is assumed to be **298K** = **25ºC**

4. The SI unit of pressure is Pascal(**Pa**) / Newton per metre squared (**Nm^{-2}**) . Millimeters' of mercury(**mmHg**) ,centimeters of mercury(**cmHg**) and **atmospheres** are also commonly used.

The units are **not** interconvertible but Pascals(Pa) are equal to Newton per metre squared(Nm^{-2}).
The **standard** pressure is the **atmospheric** pressure.
Atmospheric pressure is **equal** to about:
(i)101325 Pa
(ii)101325 Nm^{-2}
(iii)760 mmHg
(iv)76 cmHg
(v)one atmosphere.

5. Molecules of gases are always in continuous random motion at high speed. This motion is affected by the physical conditions of temperature and pressure.
Physical conditions change the volume occupied by gases in a **closed** system.
The effect of physical conditions of temperature and pressure was investigated and expressed in both Boyles and Charles laws.

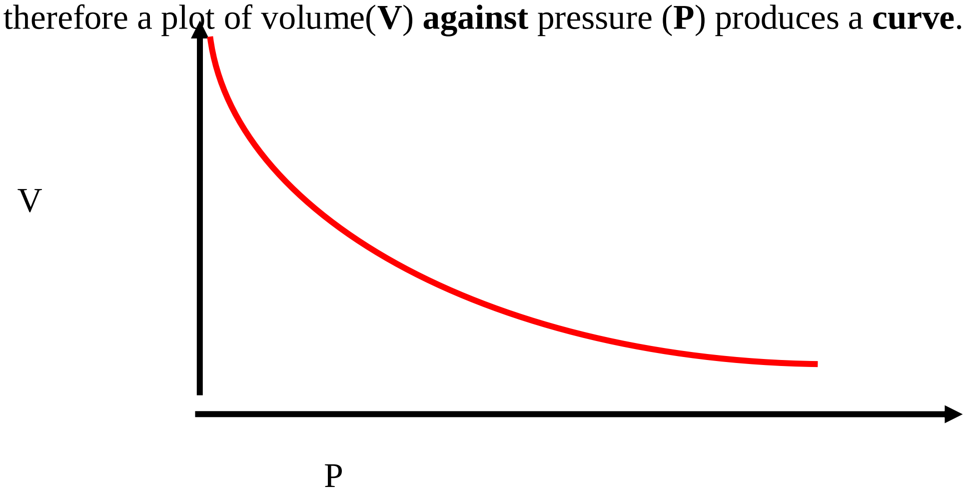
6. Boyles law states that
"**the volume of a fixed mass of a gas is inversely proportional to the pressure at constant/fixed temperature** "
Mathematically:

$$\text{Volume} \propto \frac{1}{\text{Pressure}} \quad \textbf{(Fixed /constant Temperature)}$$

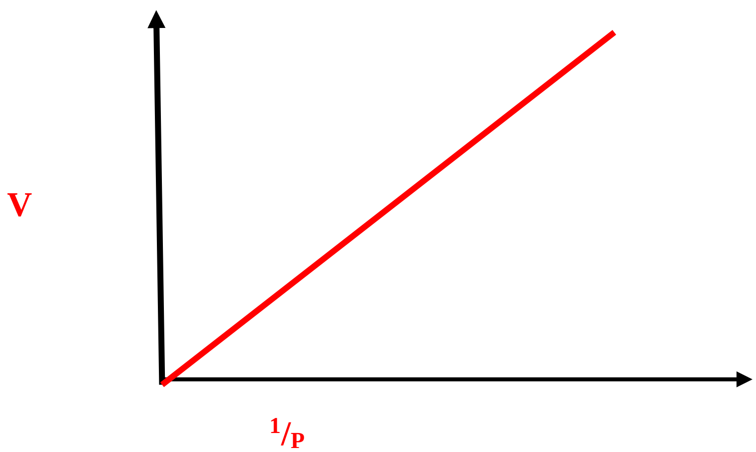
$$V \propto \frac{1}{P} \quad \textbf{(Fixed /constant T)} \text{ ie } \mathbf{PV = Constant}(k)$$

From Boyles law , an **increase** in pressure of a gas cause a **decrease** in volume. i.e **doubling** the pressure cause the volume to be **halved**.

Graphically therefore a plot of volume(**V**) **against** pressure (**P**) produces a **curve**.



Graphically a plot of volume(**V**) **against** inverse/reciprocal of pressure (**1/p**) produces a **straight line**



For **two** gases then **$P_1 V_1 = P_2 V_2$**

P_1 = Pressure of gas 1

V_1 = Volume of gas 1

P_2 = Pressure of gas 2

V_2 = Volume of gas 2

Practice examples:

1. A fixed mass of gas at 102300Pa pressure has a volume of 25cm3.Calculate its volume if the pressure is doubled.

Working

$P_1 V_1 = P_2 V_2$ Substituting :102300 x 25 = (102300 x **2**) x V_2

$$V_2 = \frac{102300 \times 25}{(102300 \times 2)} = \mathbf{12.5cm3}$$

2. Calculate the pressure which must be applied to a fixed mass of 100cm3 of Oxygen for its volume to triple at 100000Nm^{-2}.

$P_1 V_1 = P_2 V_2$ Substituting :100000 x 100 = P_2 x (100 x **3**)

$$V_2 = \frac{100000 \times 100}{(100 \times 3)} = \mathbf{33333.3333\ Nm^{-2}}$$

3.A 60cm3 weather ballon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the ballon reach stratosphere where the pressure is 90000Pa?

$P_1 V_1 = P_2 V_2$ Substituting :101325 x 60 = 90000 x V_2

$V_2 = \frac{101325 \times 60}{90000}$ = **67.55 cm3**

The new volume at 67.55 cm3 **exceed** ballon capacity of 60.00 cm3.It will **burst** before reaching destination.

7.Charles law states that"**the volume of a fixed mass of a gas is directly proportional to the absolute temperature at constant/fixed pressure** "

Mathematically:

Volume α Pressure **(Fixed /constant pressure)**

V α T **(Fixed /constant P)** ie $\frac{V}{T}$ **= Constant**(k)

From Charles law , an **increase** in temperature of a gas cause an **increase** in volume. i.e **doubling** the temperature cause the volume to be **doubled**.

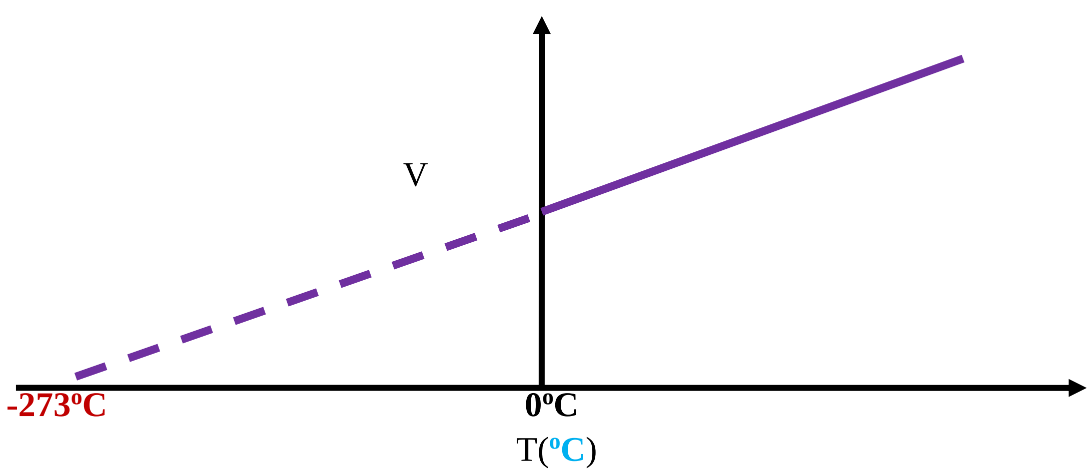
Gases expand/increase by $^1/_{273}$ by volume on heating.Gases contact/decrease by $^1/_{273}$ by volume on cooling at constant/fixed pressure.

The volume of a gas continue decreasing with decrease in temperature until at **-273°C /0 K** the volume is **zero**. i.e. there is no gas.

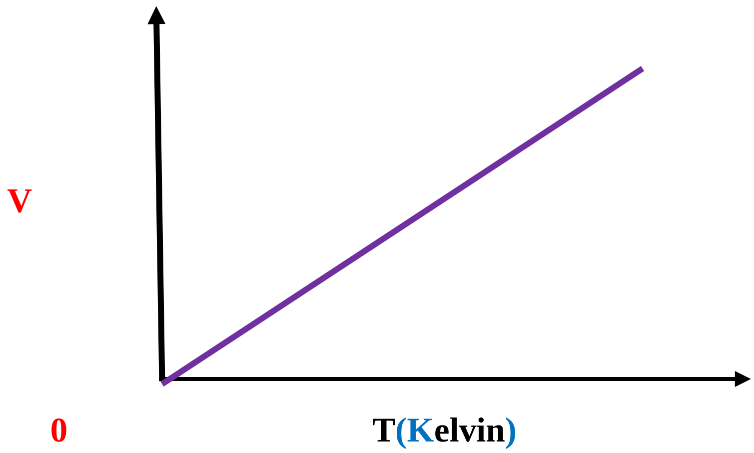
This temperature is called **absolute zero.** It is the **lowest** temperature at which a gas **can** exist.

Graphically therefore a plot of volume(**V**) **against** Temperature(**T**) in:

(**i**)°C produces a **straight line** that is **extrapolated** to the absolute zero of -273**°C** .



(**ii**)Kelvin/K produces a **straight line** from absolute zero of **O Kelvin**



For **two** gases then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

T_1 = Temperature **in Kelvin** of gas 1

V_1 = Volume of gas 1

T_2 = Temperature **in Kelvin** of gas 2

V_2 = Volume of gas 2

Practice examples:

1. 500cm3 of carbon(IV)oxide at 0°C was transfered into a cylinder at -4°C. If the capacity of the cylinder is 450 cm3,explain what happened.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ substituting $\frac{500}{(0 + 273)} = \frac{V_2}{(-4 + 273)}$

$= \frac{500 \text{ x } (-4 \text{ x } 273)}{(0 + 273)} =$ **492.674cm3**

The capacity of cylinder (500cm3) is **less** than new volume(492.674cm3).

7.326cm3(500-492.674cm3)of carbon(IV)oxide gas did not fit into the cylinder.

2. A mechanic was filling a deflated tyre with air in his closed garage using a hand pump. The capacity of the tyre was 40,000cm3 at room temperature. He rolled the tyre into the car outside. The temperature outside was 30°C.Explain what happens.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ substituting $\frac{40000}{(\mathbf{25} + 273)} = \frac{V_2}{(\mathbf{30} + 273)}$

$= \frac{40000 \text{ x } (30 \text{ x } 273)}{(25 + 273)} =$ **40671.1409cm3**

The capacity of a tyre (40000cm3) is **less** than new volume(40671.1409cm3).

The tyre thus bursts.

3. A hydrogen gas balloon with 80cm3 was released from a research station at room temperature. If the temperature of the highest point it rose is -30ºC , explain what happened.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{substituting} \quad \frac{80}{(\mathbf{25}+273)} = \frac{V_2}{(\mathbf{-30}+273)}$$

$$= \frac{80 \times (-30 \times 273)}{(25+273)} = \mathbf{65.2349cm3}$$

The capacity of balloon (80cm3) is **more** than new volume (65.2349cm3).

The balloon thus remained intact.

8. The continuous random motion of gases differ from gas to the other.The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion.**

The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas**.

The density of a gas depends on its molar mass/relative molecular mass. i.e. **The** higher **the density the** higher **the molar mass/relative atomic mass and thus the** lower **the rate of diffusion.**

Examples

1.Carbon (IV)oxide(CO_2) has a molar mass of 44g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus lighter/less dense than Carbon (IV)oxide(CO_2). N_2 diffuses faster than CO_2.

2.Ammonia(NH_3) has a molar mass of 17g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus about **twice** lighter/less dense than Ammonia(NH_3). Ammonia(NH_3) diffuses twice faster than N_2.

3. Ammonia(NH_3) has a molar mass of 17g.Hydrogen chloride gas has a molar mass of 36.5g.Both gases on contact react to form **white fumes** of ammonium chloride .When a glass/cotton wool dipped in ammonia and another glass/cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to glass/cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower .

Diffusion of ammonia and hydrogen chloride

Chemical equation: $HCl(g) + NH_3(s) \rightarrow NH_3Cl(s)$

The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**. Grahams law states that:

"the rate of diffusion of a gas is inversely proportional to the square root of its density, at the same/constant/fixed temperature and pressure"

Mathematically

$R \propto \frac{1}{\sqrt{p}}$ and since density is proportional to mass then $R \propto \frac{1}{\sqrt{m}}$

For two gases then:

$\frac{R_1}{\sqrt{M_2}} = \frac{R_2}{\sqrt{M_1}}$ where: R_1 and R_2 is the **rate** of diffusion of **1**st and **2**nd gas.
M_1 and M_2 is the **molar mass** of **1**st and **2**nd gas.

Since rate is inverse of time. i.e. the higher the rate the less the time:

For two gases then:

$\frac{T_1}{\sqrt{M_1}} = \frac{T_2}{\sqrt{M_2}}$ where: T_1 and T_2 is the **time taken** for **1**st and **2**nd gas to diffuse.
M_1 and M_2 is the **molar mass** of **1**st and **2**nd gas.

Practice examples:

1. It takes 30 seconds for 100cm3 of carbon(IV)oxide to diffuse across a porous plate. How long will it take 150cm3 of nitrogen(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,N=14.0=16.0)

Molar mass CO_2=44.0   Molar mass NO_2=46.0

Method 1

100cm3 CO_2 takes 30seconds

150cm3 takes $\frac{150 \times 30}{100}$ = 45seconds

$\frac{T\ CO_2}{T\ NO_2} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } NO_2}} \Rightarrow \frac{45\text{seconds}}{T\ NO_2} = \frac{\sqrt{44.0}}{\sqrt{46.0}}$

$T\ NO_2 = \frac{45\text{seconds} \times \sqrt{46.0}}{\sqrt{44.0}}$ = **46.0114** seconds

Method 2

100cm3 CO_2 takes 30seconds

1cm3 takes $\frac{100 \times 1}{30}$ = **3.3333cm3sec^{-1}**

$\frac{R\ CO_2}{R\ NO_2} = \frac{\sqrt{\text{molar mass } NO_2}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{3.3333\text{cm3sec}^{-1}}{R\ NO_2} = \frac{\sqrt{46.0}}{\sqrt{44.0}}$

$R\ NO_2 = 3.3333\text{cm3sec}^{-1} \times \sqrt{44.0}$ = **3.2601**cm3sec^{-1}

√ 46.0

3.2601cm3 takes 1seconds

150cm3 take $\frac{150cm3}{3.2601cm3}$ = **46.0109seconds**

2. How long would 200cm3 of Hydrogen chloride take to diffuse through a porous plug if carbon(IV)oxide takes 200seconds to diffuse through.

Molar mass CO_2 = 44g   Molar mass HCl = 36.5g

$\frac{T\ CO_2}{T\ HCl} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass HCl}}}$ => $\frac{200 \text{ seconds}}{T\ HCl} = \frac{\sqrt{44.0}}{\sqrt{36.5}}$

T HCl = $\frac{200\text{seconds} \times \sqrt{36.5}}{\sqrt{44.0}}$ = **182.1588** seconds

3. Oxygen gas takes 250 seconds to diffuse through a porous diaphragm. Calculate the molar mass of gas Z which takes 227 second to diffuse.

Molar mass O_2 = 32g Molar mass Z = x g

$\frac{T\ O_2}{T\ Z} = \frac{\sqrt{\text{molar mass } O_2}}{\sqrt{\text{molar mass } Z}}$ => $\frac{250 \text{ seconds}}{227\text{seconds}} = \frac{\sqrt{32.0}}{\sqrt{x}}$

$\sqrt{x}$ = $\frac{227\text{seconds} \times \sqrt{32}}{250}$ = **26.3828** grams

4. 25cm3 of carbon(II)oxide diffuses across a porous plate in 25seconds. How long will it take 75cm3 of Carbon(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,0=16.0)

Molar mass CO_2 = 44.0   Molar mass CO = 28.0

Method 1

25cm3 CO takes 25seconds

75cm3 takes $\frac{75 \times 25}{25}$ = 75seconds

$\frac{T\ CO_2}{T\ CO} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass CO}}}$ => $\frac{T\ CO_2\text{seconds}}{75} = \frac{\sqrt{44.0}}{\sqrt{28.0}}$

T CO_2 = $\frac{75\text{seconds} \times \sqrt{44.0}}{\sqrt{28.0}}$ = **94.0175** seconds

Method 2

25cm3 CO_2 takes 25seconds

1cm3 takes $\frac{25 \times 1}{25}$ = **1.0cm3sec^{-1}**

$$\frac{R\ CO_2}{R\ CO} = \frac{\sqrt{\text{molar mass CO}}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{x\ \text{cm3sec}^{-1}}{1.0\text{cm3sec}^{-1}} = \frac{\sqrt{28.0}}{\sqrt{44.0}}$$

$$R\ CO_2 = \frac{1.0\text{cm3sec}^{-1} \times \sqrt{28.0}}{\sqrt{44.0}} = \mathbf{0.7977}\text{cm3sec}^{-1}$$

0.7977cm3 takes 1 seconds

75cm3 takes $\frac{75\text{cm3}}{0.7977\text{cm3}}$ = **94.0203seconds**

THE MOLE

THE MOLE-FORMULAE AND CHEMICAL EQUATIONS (40 LESSONS)

Introduction to the mole, molar masses and Relative atomic masses

1. The mole is the **SI** unit of the **amount** of substance.
2. The number of particles e.g. atoms, ions, molecules, electrons, cows, cars are all measured in terms of moles.
3. The number of particles in one mole is called the **Avogadros Constant**. It is denoted "**L**".

The Avogadros Constant contain **6.023 x10 ^{23}** particles. i.e.

1mole = 6.023 x10 ^{23} particles = 6.023 x10 ^{23}
2 moles = 2 x 6.023 x10 ^{23} particles = 1.205 x10 ^{24}
0.2 moles = 0.2 x 6.023 x10 23 particles = 1.205 x10 ^{22}
0.0065 moles = 0.0065 x 6.023 x10 ^{23} particles = 3.914 x10 ^{21}

3. The mass of one mole of a substance is called **molar mass**. The molar mass of:

(i)an **element** has mass equal to relative **atomic** mass /RAM(in grams)of the element e.g.
Molar mass of carbon(C)= relative atomic mass = 12.0g
6.023 x10 ^{23} particles of carbon = 1 mole =12.0 g

Molar mass of sodium(Na) = relative atomic mass = 23.0g
6.023 x10 ^{23} particles of sodium = 1 mole =23.0 g

Molar mass of Iron (Fe) = relative atomic mass = 56.0g
6.023 x10 ^{23} particles of iron = 1 mole =56.0 g

(ii)a **molecule** has mass equal to relative molecular mass /RMM (in grams)of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.

The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic (e.g. O_2, H_2, N_2, F_2, Cl_2, Br_2, I_2**)noble gases are **monoatomic(e.g. He, Ar, Ne, Xe)**,Ozone gas(**O_3**) is **triatomic** e.g.

Molar mass **Oxygen molecule(O_2)** =relative molecular mass =(16.0x 2)g =32.0g
6.023 x10 ^{23} particles of Oxygen molecule = 1 mole = 32.0 g

Molar mass **chlorine molecule(Cl_2)** =relative molecular mass =(35.5x 2)g =71.0g

6.023×10^{23} particles of chlorine molecule = 1 mole = 71.0 g

Molar mass **Nitrogen molecule(N_2)** =relative molecular mass =(14.0x 2)g =28.0g
6.023×10^{23} particles of Nitrogen molecule = 1 mole = 28.0 g

(ii)a **<u>compound</u>** has mass equal to <u>relative formular mass</u> /RFM (in grams)of the molecule. Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.

(i)Molar mass **Water(H_2O)** = relative formular mass =[(1.0 x 2) + 16.0]g =18.0g

6.023×10^{23} particles of Water molecule = 1 mole = 18.0 g
6.023×10^{23} particles of Water molecule has:
- 2 x 6.023×10^{23} particles of Hydrogen atoms
-1 x 6.023×10^{23} particles of Oxygen atoms

(ii)Molar mass **sulphuric(VI)acid(H_2SO_4)** = relative formular mass
=[(1.0 x 2) + 32.0 + (16.0 x 4)]g =98.0g

6.023×10^{23} particles of sulphuric(VI)acid(H_2SO_4) = 1 mole = 98.0g
6.023×10^{23} particles of sulphuric(VI)acid(H_2SO_4) has:
- 2 x 6.023×10^{23} particles of **H**ydrogen atoms
-1 x 6.023×10^{23} particles of **S**ulphur atoms
-4 x 6.023×10^{23} particles of **O**xygen atoms

(iii)Molar mass **sodium carbonate(IV)(Na_2CO_3)** = relative formular mass
=[(23.0 x 2) + 12.0 + (16.0 x 3)]g =106.0g

6.023×10^{23} particles of sodium carbonate(IV)(Na_2CO_3) = 1 mole = 106.0g
6.023×10^{23} particles of sodium carbonate(IV)(Na_2CO_3) has:
- 2 x 6.023×10^{23} particles of **S**odium atoms
-1 x 6.023×10^{23} particles of **C**arbon atoms
-3 x 6.023×10^{23} particles of **O**xygen atoms

(iv)Molar mass **Calcium carbonate(IV)($CaCO_3$)** = relative formular mass
=[(40.0+ 12.0 + (16.0 x 3)]g =100.0g.

6.023×10^{23} particles of Calcium carbonate(IV)($CaCO_3$) = 1 mole = 100.0g
6.023×10^{23} particles of Calcium carbonate(IV)($CaCO_3$) has:
- 1 x 6.023×10^{23} particles of **C**alcium atoms
-1 x 6.023×10^{23} particles of **C**arbon atoms
-3 x 6.023×10^{23} particles of **O**xygen atoms
(v)Molar mass **Water(H_2O)** = relative formular mass

=[(2 x 1.0)+ 16.0]g =18.0g

6.023 x10 ^{23} particles of Water(H_2O) = 1 mole = 18.0g
6.023 x10 ^{23} particles of Water(H_2O) has:
- 2 x 6.023 x10 ^{23} particles of Hydrogen atoms
-2 x 6.023 x10 ^{23} particles of **O**xygen atoms

<u>Practice</u>

1. Calculate the number of moles present in:

(i)0.23 g of Sodium atoms
Molar mass of Sodium atoms = 23g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23g}{23} = \mathbf{0.01moles}$$

(ii) 0.23 g of Chlorine atoms
Molar mass of Chlorine atoms = 35.5 g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23g}{35.5} = \mathbf{0.0065moles\ /6.5 \times 10^{-3}\ moles}$$

(iii) 0.23 g of Chlorine molecules
Molar mass of Chlorine molecules =(35.5 x 2) = **71.0** g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23g}{71} = \mathbf{0.0032moles\ /3.2 \times 10^{-3}\ moles}$$

(iv) 0.23 g of dilute sulphuric(VI)acid
Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23g}{98} = \mathbf{0.0023moles\ /2.3 \times 10^{-3}\ moles}$$

2. Calculate the number of atoms present in:(Avogadros constant L = 6.0 x 10 ^{23})

(i) 0.23 g of dilute sulphuric (VI)acid

<u>Method I</u>

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23g}{98} = \mathbf{0.0023moles\ /2.3 \times 10^{-3}\ moles}$$

1 mole has 6.0 x 10 ^{23} atoms

$$2.3 \times 10^{-3} \text{ moles has } \left(\frac{2.3 \times 10^{-3} \times 6.0 \times 10^{23}}{1}\right) = \mathbf{1.38 \times 10^{21}} \text{ atoms}$$

<u>Method II</u>

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g
98.0g = 1 mole has 6.0 x 10 ^{23} atoms

$$0.23 \text{ g therefore has } \left(\frac{0.23 \text{ g} \times 6.0 \times 10^{23}}{98}\right) = \mathbf{1.38 \times 10^{21}} \text{ atoms}$$

(ii)0.23 g of sodium carbonate(IV)decahydrate

Molar mass of $Na_2CO_3.10H_2O$=

[(2 x 23) + 12 + (3 x16) + (10 x 1.0) + (10 x 16)] = **276.0**g

Method I

Moles = $\frac{\text{mass in grams}}{\text{Molar mass}}$ => $\frac{0.23g}{276}$ = **0.00083moles / 8.3 x 10^{-4} moles**

1 mole has 6.0 x 10^{23} atoms

8.3 x 10^{-4} moles has $\frac{(8.3 \times 10^{-4} \text{ moles} \times 6.0 \times 10^{23})}{1}$ = **4.98 x 10^{20}** atoms

Method II

276.0g = 1 mole has 6.0 x 10^{23} atoms

0.23 g therefore has $\frac{(0.23 \text{ g} \times 6.0 \times 10^{23})}{276.0}$ = **4.98 x 10^{20}** atoms

(iii)0.23 g of Oxygen gas

Molar mass of O_2 = (2 x16) = **32.0** g

Method I

Moles = $\frac{\text{mass in grams}}{\text{Molar mass}}$ => $\frac{0.23g}{32}$ = **0.00718moles / 7.18 x 10^{-3} moles**

1 mole has **2** x 6.0 x 10^{23} **atoms** in O_2

7.18 x 10^{-3}moles has $\frac{(7.18 \times 10^{-3}\text{moles} \times 2 \times 6.0 \times 10^{23})}{1}$ =**8.616 x 10^{21}**atoms

Method II

32.0g = 1 mole has **2** x 6.0 x 10^{23} **atoms** in O_2

0.23 g therefore has $\frac{(0.23 \text{ g} \times 2 \times 6.0 \times 10^{23})}{32.0}$ = **8.616 x 10^{21}**atoms

(iv)0.23 g of Carbon(IV)oxide gas

Molar mass of CO_2 = [12 + (2 x16)] = **44.0** g

Method I

Moles = $\frac{\text{mass in grams}}{\text{Molar mass}}$ => $\frac{0.23g}{44}$ = **0.00522moles / 5.22 x 10^{-3} moles**

1 mole has **3** x 6.0 x 10^{23} **atoms** in CO_2

7.18 x 10^{-3}moles has $\frac{(5.22 \times 10^{-3}\text{moles} \times 3 \times 6.0 \times 10^{23})}{1}$ =**9.396 x 10^{21}**atoms

Method II

44.0g = 1 mole has **3** x 6.0 x 10^{23} **atoms** in CO_2

0.23 g therefore has $\left(\frac{0.23\text{ g} \times 3 \times 6.0 \times 10^{23}}{44.0}\right)$ = **9.409 x 10^{21}**atoms

Empirical and molecular formula

NOTE!

This is a Sample of the Well Organized Detailed Simplified Notes Available.

Call/Text/WhatsApp 0746-222-000 for the Complete Notes.

FOR THE FOLLOWING;

- ✓ **ONLINE TUITION**
- ✓ **REVISION NOTES**
- ✓ **SCHEMES OF WORK**
- ✓ **SETBOOKS VIDEOS**
- ✓ **TERMLY EXAMS**
- ✓ **QUICK REVISION KITS**
- ✓ **KCSE TOPICALS**
- ✓ **KCSE PREMOCKS**
- ✓ **TOP SCHOOLS PREMOCKS**
- ✓ **JOINT PREMOCKS**
- ✓ **KCSE MOCKS**
- ✓ **TOP SCHOOLS MOCKS**
- ✓ **JOINT MOCKS**
- ✓ **KCSE POSTMOCKS**

- ✓ **TOP SCHOOLS PREDICTIONS**
- ✓ **KCSE PREDICTIONS**
- ✓ **KCSE REVEALED SETS**

CALL/TEXT 0746 222 000

mwalimuconsultancy@gmail.com

POWERED BY MR ISABOKE

SUCCESS