A. BOYLE'S LAW

- It relates the **pressure** and **volume** of a **fixed mass of gas** at **constant temperature**.
- The arrangements in **figure 10.1 (a)** and **(b)** can be used to demonstrate the relationship between pressure and volume of a fixed mass of gas at constant temperature.

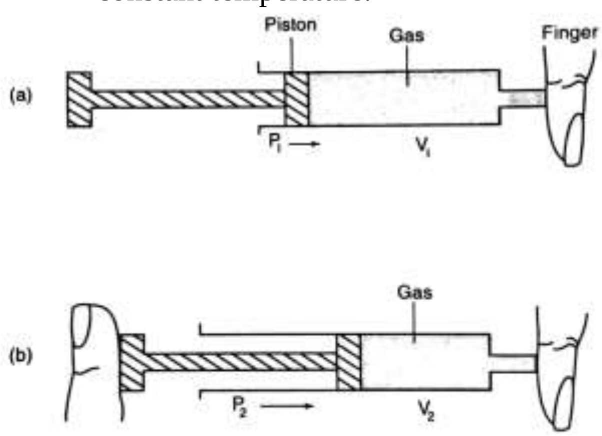


Figure 10.1

The nozzle of the syringe is closed with a finger and the piston slowly pushed inwards.

Observation
It is observed that ***an increase in pressure of a mass of gas results in decrease in volume***.

EXPERIMENT 10.1: To investigate the relationship between pressure and volume of a fixed mass of a gas at constant temperature.

Apparatus

Thick-walled J-shaped glass tube with one end closed, oil, Bourdon gauge, foot pump, metre rule.

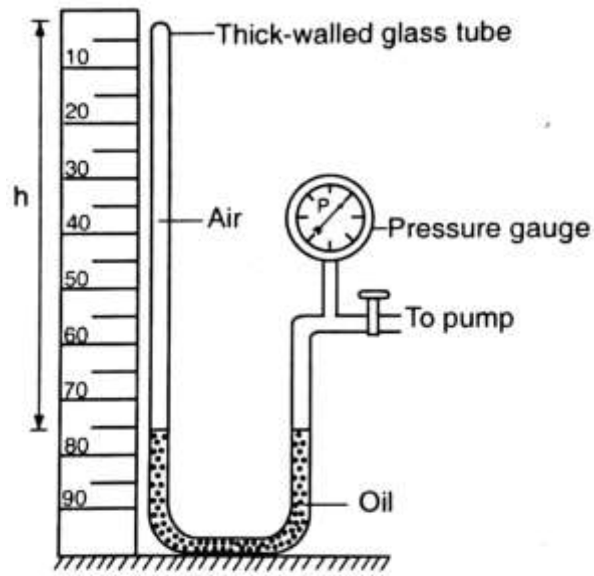


Figure 10.2

Procedure
- Set up the apparatus as shown in **figure 10.2**.
- Connect the foot pump to the apparatus and with the tap open, pump in air until the oil rises a small but measurable height, then close the tap.
- Allow the air to adjust to room temperature, then read the value of the pressure, P on the gauge and the height, h of the air column.

Note:

The **height, h of the air column** represents the **volume of the air**, since the glass tube has a ***uniform area of cross-section***. (***Volume = Area of uniform cross-section x Height***)

- Repeat the experiment by varying the values of pressure to obtain corresponding readings of height of the air column.
- Record your results in the table 1.

Table 1

Pressure, P (Pa)	**Volume, V (h cm)**	$\frac{1}{V}\left(\frac{1}{h}cm^{-1}\right)$	**PV**

- Using the results in the table, plot a graph of:

(a) **P** against **V**.

(b) **P** against $\frac{1}{V}$**.** Determine the slope.

(c) **PV** against **P**.

Results and Conclusion

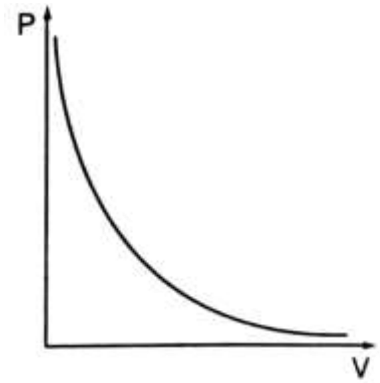
- This experiment shows that ***an increase in pressure of a fixed mass of gas causes a decrease in its volume.***
- This is summarized in **Boyle's Law**.
- It states that:

The pressure of a fixed mass of a gas is inversely proportional to its volume, provided that the temperature is kept constant.

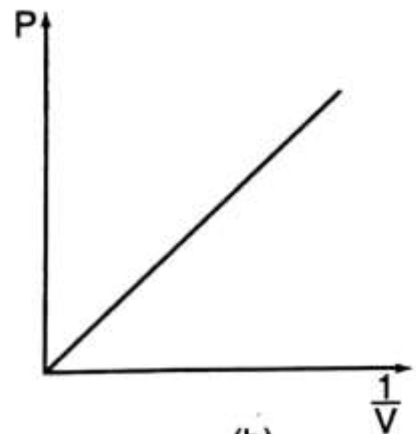
- In symbols;

$\mathbf{P}\ \frac{1}{V}$, or, $\mathbf{P} = k\ x\ \frac{1}{V}$

- So, **PV = constant**, i.e., $P_1V_1 = P_2V_2$**,** for any given mass of a gas**.**
- **Figure 10.3 (a), (b)** and **(c)** show sketches of the relationship between pressure, P and volume, V of a fixed mass of gas at constant temperature.



(a)



(b)

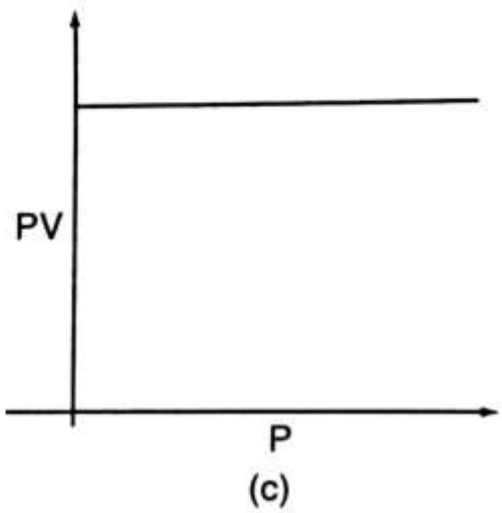


Figure 10.3

Note:

(a) If the experiment above is repeated at **different temperatures**, similar curves are obtained as shown in **figure 10.4**.

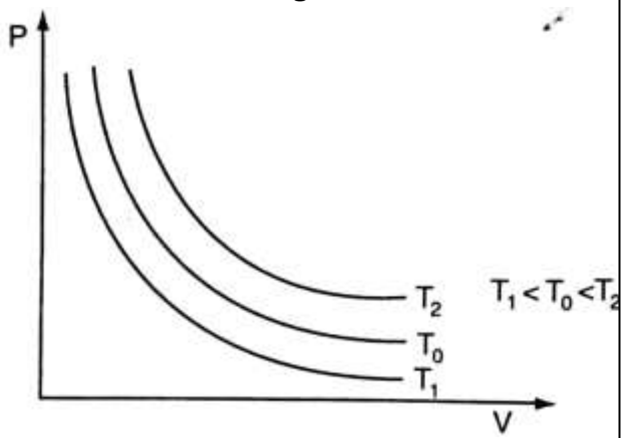


Figure 10.4

Each of the curves is called an **isothermal curve.**

(b) When **P** is plotted against $\frac{1}{V}$ for each of the isothermals, the results obtained would be as shown in **figure 10.5.**

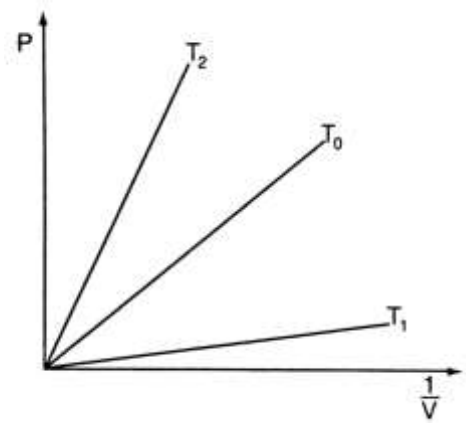


Figure 10.5

Example 1

Determine the pressure required to compress a gas in a cylinder initially at 20 ^{0}C and at a pressure of 1.03 x 10^5 Pa to one-eighth of its original volume.

Solution

$\boldsymbol{P_1V_1 = P_2V_2 \leftrightarrow P_2 = \frac{P_1V_1}{V_2}}$

$\boldsymbol{P_1 = 1.03 \times 10^5}$ **Pa**, $\boldsymbol{V_1 = v, V_2 = \frac{1}{8}v}$

$\boldsymbol{P_2} = \frac{1.03 \times 10^5 \times V}{\frac{1}{8}V}$

$\underline{P_2 = 8.24 \times 10^5 \text{ Pa}}$

Example 2

Figure 10.6 shows an air bubble of volume 2.0 cm^3 at the bottom of a lake 40 m deep.

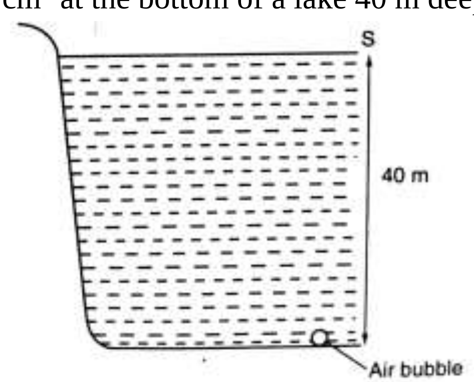


Figure 10.6

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

Solution

10 m water column = 1 atmosphere

Therefore, 40 m water column = 4 atmospheres.

Total pressure, P_1 at the bottom = 1 + 4 = 5 atmospheres

Pressure P_2 at the surface = 1 atm.

Volume, V_1 at the bottom = 2.0 cm^3

By Boyle's law, $P_1V_1 = P_2V_2$

5 x 2.0 = 1 x V_2

V_2 = 10.0 cm^3

Example 3

A column of air 26 cm long is trapped by a mercury thread 5 cm long as shown in **figure 10.7 (a)**. When the tube is inverted as in **figure 10.7 (b)**, the air column becomes 30 cm long.

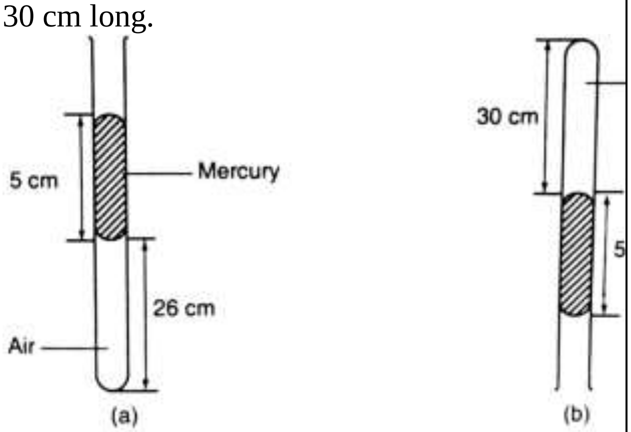


Figure 10.7

Determine the value of the atmospheric pressure.

Solution

Volume of the air column is proportional to length of the air column

Form Boyle's Law, $P_1V_1 = P_2V_2$

In fig (a), air pressure = atm. pressure + ρgh

In fig (b), air pressure = atm. pressure - ρgh

ρ = Density of mercury.

Let the atm. pressure be x cm of mercury.

(x + 5) x 26 = (x- 5) x 30

26x + 130 = 30x – 150

280 = 4x

x = 70 cmHg

Example 4

Figure 10.8 shows a graph of pressure P, against volume, V, for a fixed mass of gas at constant temperature.

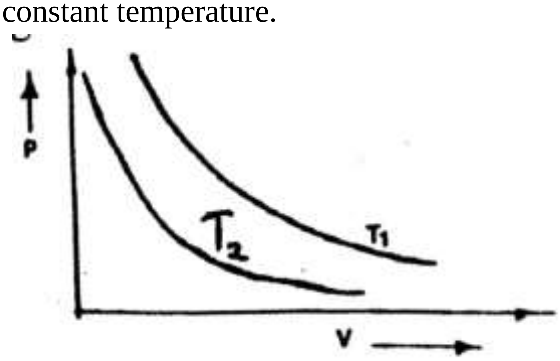


Figure 10.8

Sketch on the same axes, a graph for the same mass at gas with a temperature T_2 lower than T_1. (1 mark)

Example 5

Figure 10.9 shows a set-up that may be used to verify Boyle's law.

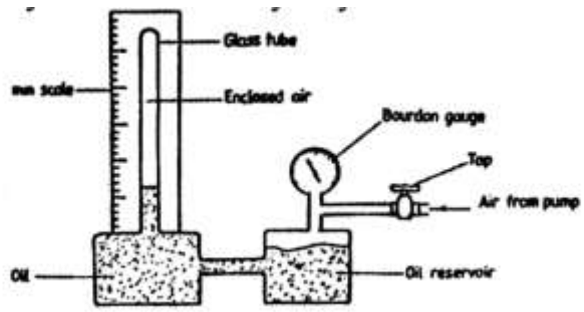


Figure 10.9

(a) Describe the measurements that should be taken in the experiment. (2 marks)

(i) Pressure by use of bourdon gauge.

(ii) Length (or volume) of the enclosed air column.

(b) Explain how the measurements taken in (a) above may be used to verify Boyle's law. (4 marks)

- **Adjust the pressure entering from the pump and note the corresponding length (or volume) of the enclosed air.**
- **Repeat the experiment for other values of the pressure and tabulate the values.**
- **Plot the graph of $\frac{1}{V}$ against P. The graph is a straight line passing through the origin, showing that $P \propto \frac{1}{V}$.**

Example 6

Two identical containers A and B are placed on a bench. Container A is filled with oxygen gas and container B with hydrogen gas such that the two gases have equal masses. If the containers are maintained at the same temperature, state with reason the container in which the pressure is higher. (2 marks)

The pressure in B is higher than in A. There are more hydrogen gas molecules than oxygen gas molecules. The collision of hydrogen gas molecules with the walls of the container is therefore higher in B.

B. CHARLES' LAW

- It relates **volume** and **temperature of a fixed mass** of gas at **constant pressure**.
- The set up in **figure 10.10** can be used to demonstrate the relationship between temperature and volume of a fixed mass of a gas at constant pressure.

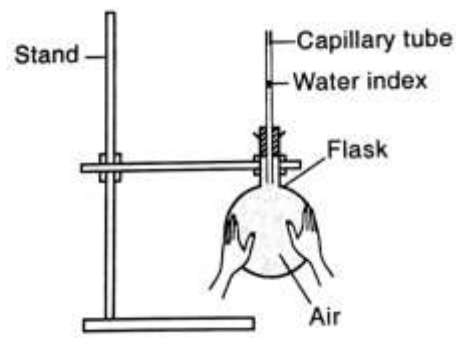


Figure 10.10

The flask is grasped firmly and the water index observed.

Observation

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.

EXPERIMENT 2: To Investigate the Relationship between Volume and Temperature of a Fixed Mass of a Gas at Constant Pressure.

Apparatus

Capillary tube sealed at one end, concentrated sulphuric acid, thermometer, half metre rule, stirrer, source of heat, retort stand, rubber band, water bath.

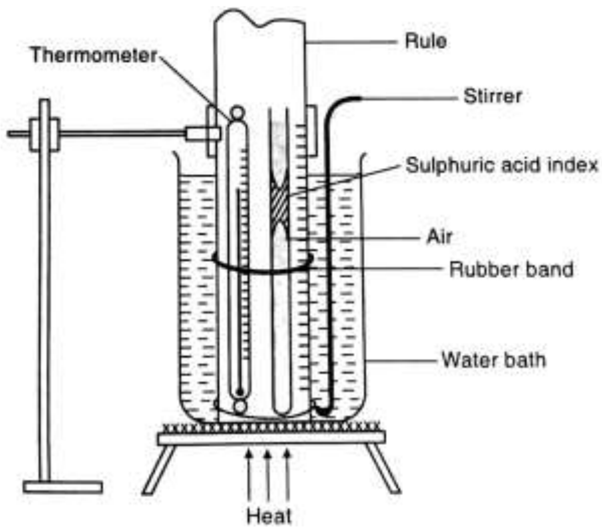


Figure 10 .11

Procedure

- Introduce concentrated sulphuric acid deep into the glass tube to trap air in the tube.
- Attach the tube, thermometer and the half metre rule using the rubber band.
- Assemble the apparatus as shown in the **figure 10.11**.
- Record the room temperature and the corresponding height, h of air column in the tube.
- Heat the water bath and record the temperature and the corresponding height at suitable temperature intervals in the table 2.

Table 2

Temperature (θ ^{0}C)						
Height, h (cm)						

Note:
(i) The sulphuric acid index serves as a pointer to the volume (height) of the gas on the scale as well as a drying agent for the air.
(ii) Pressure of the trapped air is equal to the atmospheric pressure plus pressure due to the sulphuric acid index, which remains constant throughout the experiment.

(iii) Before taking the readings, stir the water bathe so that the temperature of the gas is equal to that of the water bath.

Observation

- As the temperature rises, the height, h (volume) of the gas also increases.
- **Figure 10.12** shows a graph of volume (cm^3) against temperature (^{0}C).

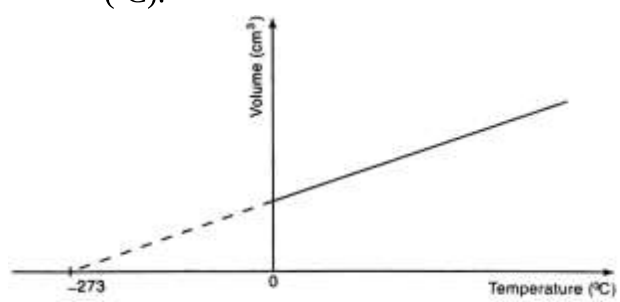


Figure 10.12

- The graph is a straight line, indicating proportional changes in volume and temperature.
- The graph cuts the volume axis above the origin.

- If the graph is extrapolated, it cuts the temperature axis at about $^-273\ ^0C$.
- At the temperature of $^-273\ ^0C$, the volume of a gas is assumed to be zero.
- $^-273\ ^0C$ is the lowest possible temperature that a gas can fall to.
- $^-273\ ^0C$ is called the **absolute zero**.
- The scale of temperature based on the **absolute zero** is called the **absolute scale** or **Kelvin scale** of temperature.
- A graph of volume against absolute temperature is a straight line that passes through the origin, **figure 10.13.**

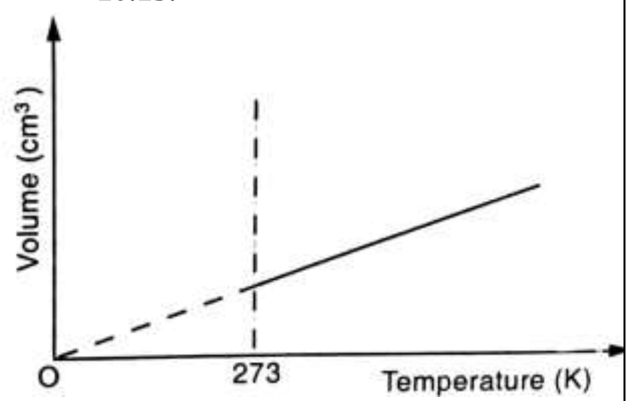


Figure 10.13

Note:
It is impossible to get to absolute zero for gases because they *condense at fairly higher temperatures.*

- It follows that on the Kelvin scale, the volume of the gas is directly proportional to the absolute (or Kelvin) temperature.
- This relation is called **Charles' Law**. It states that:

The volume of a fixed mass of a gas is directly proportional to its absolute temperature, provided that the pressure is kept constant.

- In symbols:

V α T or **V = kT** and $\mathbf{k} = \frac{V}{T}$ where k is a constant of proportionality**.**

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

Note:

This formula is only applicable when the **temperature** is expressed in **Kelvin**.

<u>Relation between Celsius and Kelvin Scale</u>

Figure 10.14 relates the Kelvin (Absolute) scale to the Celsius scale.

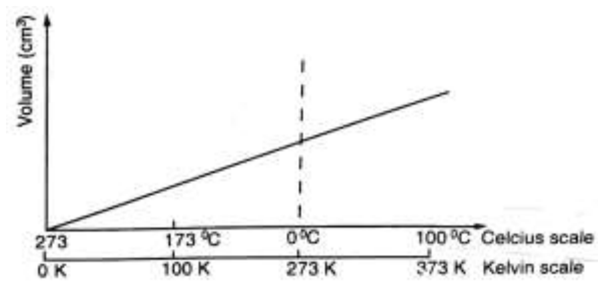


Figure 10.14

(i) To convert temperature in **degrees Celsius** to **Kelvin, add 273** to the **Cel sius** temperature i.e.

$$\boldsymbol{\theta\ ^0C = (\theta + 273)\ K}$$

(ii) To convert temperature in **Kelvin** to **degrees Celsius, subtract 273** to the **Celsius** temperature i.e.

$$\boldsymbol{\theta\ K = (\theta - 273)\ ^0C}$$

Example 7

Convert each of the following Celsius temperatures to Kelvin temperatures.

(a) $0\ ^0C = (0 + 273)\ K =$ **<u>273 K</u>**

(b) 50 ^{0}C = (50 + 273) K = **<u>323 K</u>**

(c) $^-$92 ^{0}C = ($^-$92 + 273) K = **<u>181 K</u>**

(d) 100 ^{0}C = (100 + 273) K = **<u>373 K</u>**

(e) $^-$273 ^{0}C = ($^-$273 + 273) K = **<u>0 K</u>**

Example 8

State what is meant by absolute zero temperature (Zero Kelvin 0r $^-$273 ^{0}C) (1 mark)

This is the temperature at which an ideal gas has zero volume.

Example 9

0.02 m^3 of a gas at 27 ^{0}C is heated at constant pressure until the volume is 0.03 m^3. Calculate the final temperature of the gas in ^{0}C.

Solution

$\frac{\mathbf{V}}{\mathbf{T}}$ **= constant,** i.e. $\frac{V_1}{T_1} = \frac{V_2}{T_2} \leftrightarrow T_2 = \frac{T_1 \times V_2}{V_1}$

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

T_2 = 450 K

450 K = (450 – 273) ^{0}C = <u>177 ^{0}C</u>

Example 10

A mass of air of volume 750 cm^3 is heated at constant pressure from 10^0**C** to 100 ^{0}**C.** Calculate the final volume of the air.

Solution

$\frac{V}{T}$ **= constant,** i.e. $\frac{V_1}{T_1} = \frac{V_2}{T_2} \leftrightarrow V_2 = \frac{T_2 \times V_1}{T_1}$

$V_2 = \frac{373 \times 750}{283}$

<u>V_2 = 988.5 m^3</u>

Example 11

Figure 10.15 shows a set-up that may be used to verify Charles' law.

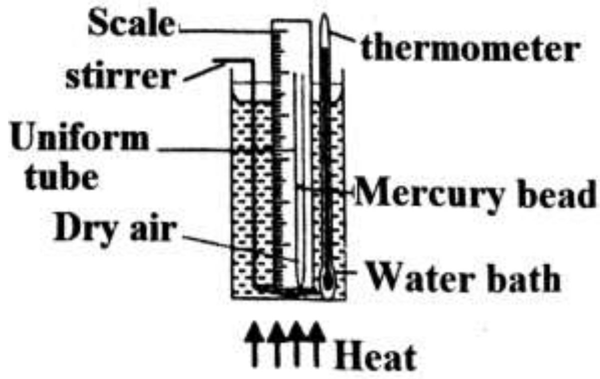


Figure 10.15

(a) State the measurements that should be taken in the experiment. (2 marks)

(i) Length (or volume) of the air column.

(ii) Temperature

(b) Explain how the measurements taken in (a) above may be used to verify Charles' law. (4 marks)

- **Air is trapped by thin mercury thread in a capillary tube.**
- **The initial temperature is noted.**
- **The water in the bath is heated and this in turn heats up the air inside the capillary tube.**
- **The length (or volume) of the air column is read against the scale and the temperature is recorded on the thermometer.**
- **Several values of temperature and the corresponding length of the air column are recorded.**
- **A graph of volume of air against temperature is drawn. The graph is a straight line cutting the temperature axis at $^-$273 ^{0}C.**

(c) What is the purpose of the water bath? (1 mark)

It allows the air to be heated uniformly.

Example 12

State the law that relates the volume of a gas to the temperature of a gas. (1 mark)

Charles' law: It states that, for a fixed mass of gas at constant pressure, the volume is directly proportional to the absolute temperature.

Example 13

A long 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 17 ^{0}C. Determine the length of the air column at 25 ^{0}C. (3 marks)

Since the area of cross-section of the bore is uniform, then the length of air column is directly proportional to volume.

V_1 = 142 mm, T_1 = 17 + 273 = 290 K,

V_2 =?, T_2 = 25 + 273 = 298 K

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ **(Charles' law)**

$\frac{142}{290} = \frac{V_2}{298}$

$V_2 = \frac{142 \times 298}{290}$

<u>**V_2 = 145.92 mm**</u>

C. PRESSURE LAW

This law relates pressure of a fixed mass of gas to its absolute temperature at constant volume.

EXPERIMENT 3: To investigate the relationship between pressure and temperature of a fixed mass of a gas at constant volume.

Apparatus

Round-bottomed flask, tight-fitting rubber cork with two holes, pressure gauge, water bath, Bunsen burner, tripod stand, thermometer, retort stand.

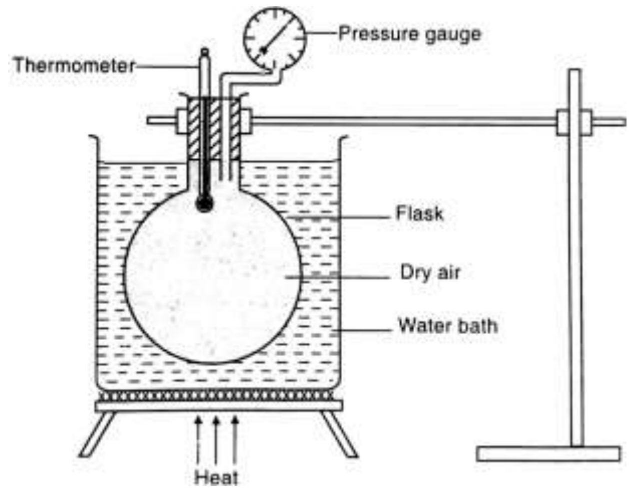


Figure 10.16

Procedure

- Set up the apparatus as shown in the **figure 10.16**.
- Record the initial temperature and pressure readings.
- Heat the water bath gently and obtain at least seven more pairs of readings at suitable temperature intervals.
- Record your results in the table 3.

Table 3

Temperature, T (^{0}C)							
Pressure, P (Pa)							
$\boldsymbol{\frac{P}{T}}$							

N*ote:*

The air in the tube connecting the pressure gauge to the glass flask may be at a lower temperature than the air in the flask. This tube should therefore be as short as possible.

Observation

- Increase in temperature causes an increase in pressure.
- The graph of pressure against time is a straight line that cuts the pressure axis above the origin, **figure 10.17**.

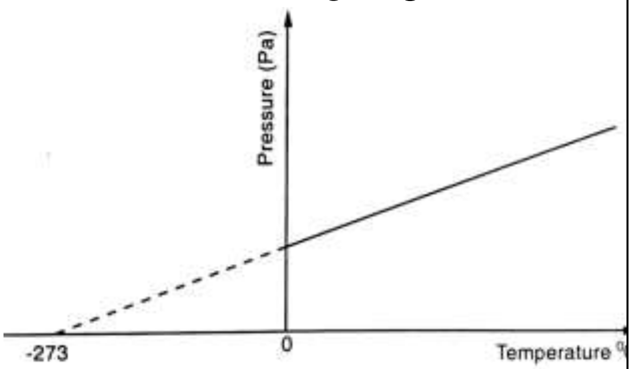


Figure 10.17

- When the graph is extrapolated, it cuts the temperature axis at ⁻273 ^{0}C, the absolute zero.
- A graph of pressure against absolute temperature is a straight line that passes through the origin, **figure 10.18**.

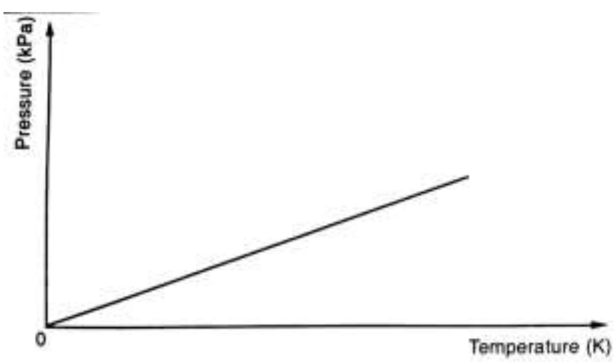


Figure 10.18

Conclusion

- On the absolute scale, the pressure of a gas is directly proportional to its absolute temperature.
- This conclusion is summed up in **pressure law**. It states that:

The pressure of a fixed mass of gas is directly proportional to its absolute temperature, provided that volume is kept constant.

In symbols:

$\mathbf{P \; \alpha \; T}$ or $\mathbf{P = kT} \leftrightarrow \mathbf{k = \frac{P}{T}}$, where k is a constant of proportionality.

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

Example 14

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

Solution

$$\frac{P_1}{T_1} = \frac{\mathrm{P_2}}{\mathrm{T_2}} \leftrightarrow \mathrm{P_2} = \frac{T_2 \times P_1}{T_1}$$

$$\mathrm{P_2} = \frac{\mathbf{373 \times 1}}{\mathbf{273}}$$

$\underline{\mathrm{P_2} = 1.366 \text{ atmospheres}}$

Example 15

At 20 ^{0}C, the pressure of a gas is 50 cm o mercury. At what temperature would the pressure of the gas fall to 10 cm of mercury?

Solution

$$\frac{\mathrm{P_1}}{\mathrm{T_1}} = \frac{P_2}{T_2} \leftrightarrow \mathrm{T_2} = \frac{T_1 \times P_2}{P_1}$$

$$\mathrm{T_2} = \frac{\mathbf{293 \times 10}}{\mathbf{50}}$$

$\underline{\mathrm{T_2} = 58.6 \text{ K (or } {}^-214.4\ {}^0\text{C)}}$

Example 16

(a) State the pressure law of an ideal gas. (1 mark)

The pressure of a fixed mass of an ideal gas is directly proportional to its absolute

temperature, provided that volume is kept constant.

(b) **Figure 10.19** shows a simple set up for pressure law apparatus.

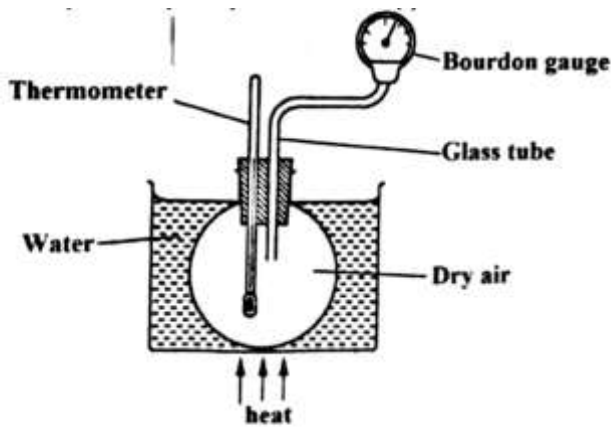


Figure 10.19

(i) (a) State the measurements that should be taken in the experiment. (2 marks)

(i) Pressure

(ii) Temperature

(b) Explain how the measurements taken in (a) above may be used to verify Charles' law. (5 marks)

- **The water in the bath is heated and the air in the flask in turn gets heated.**
- **The temperature of the gas is noted and the corresponding value of pressure noted on the bourdon gauge.**
- **Several values of the temperature, T and the corresponding values of pressure, P are tabulated.**
- **A graph of P against T is drawn,**
- **The graph is a straight line passing through the origin, indicating het pressure is directly proportional to the temperature.**

Example 17

The pressure acting on a gas in a container was changed steadily while the temperature of the gas was maintained constant. The value of the volume, **V**, of the gas was measured for various values of pressure. The graph in **figure 10.20** shows the relation between the pressure, **P**, and the reciprocal of the volume, $\frac{1}{V}$.

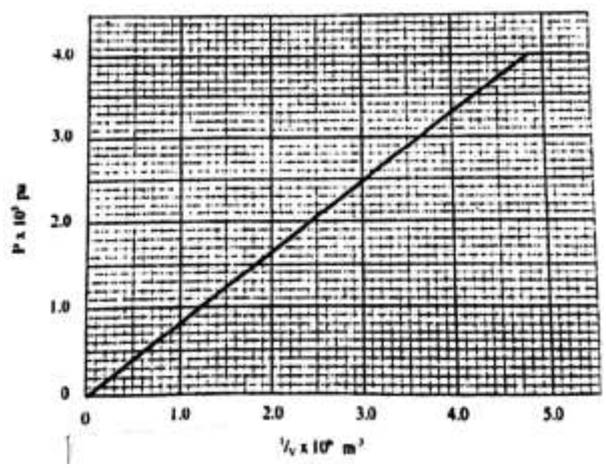


Figure 10.20

(a) Suggest how the temperature of the gas could be kept constant. (1 mark)

By changing the pressure very slowly or by allowing gas to go to original temperature after the change.

(b) 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 marks)

$$\mathbf{P} = \left(\frac{1}{V}\right)k \Rightarrow k \text{ is the slope of the graph.}$$

$$\mathbf{k} = \frac{(2.45 - 0.8) \times 10^5 \text{ Nm}^{-2}}{(3.0 - 1.0) \times 10^6 \text{ m}^{-3}}$$

k = 0.0825 Nm

(c) What physical quantity does k represent? (1 mk)

The work done on the gas

(d) State **one** precaution you would take when performing such an experiment. (1 mark)

(i) Use dry gas
(ii) Make very small changes in temperature.

Example 18

An air bubble is released at the bottom of a tall jar containing a liquid. The height of the liquid column is 80 cm. The volume of the bubble increases from 0.5 cm^3 at the bottom of the liquid to 1.15 cm^3 at the top. **Figure 10.21** shows the variation of pressure, P, on the bubble with the reciprocal of the volume, $\frac{1}{V}$, as it rises in the liquid.

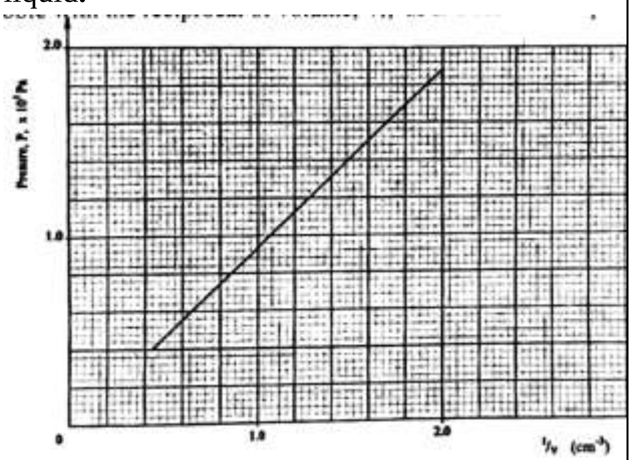


Figure 10.21

(a) State the reason why the volume increases as the bubble rises in the liquid. (1 mark)

This is because the pressure due to the liquid column decreases and therefore the pressure inside the bubble exceeds that of outside causing the bubble to expand.

(b) From the graph, determine the pressure on the bubble:

(i) At the bottom of the liquid column. (2 marks)

$\mathbf{P = \frac{1}{V}}$

$\mathbf{P = \frac{1}{0.5}}$

P = 2 Pa

The corresponding pressure is 1.88 x 10^5 Pa

(ii) At the top of the liquid column. (1 mark)

$\mathbf{P = \frac{1}{V}}$

$\mathbf{P = \frac{1}{1.15}}$

P = 0.8698 Pa

The corresponding pressure is 8.0 x 10^4 Pa

(c) Hence determine the density of the liquid in kgm^{-3}. *(Take g = 10 Nkg^{-1})* (3 marks)

Change in pressure, ΔP =
1.88 x 10^5 - 8.0 x 10^4
ΔP = 1.08 x 10^5 Pa
But $\Delta P = \rho gh$
ρ x 10 x 0.8 = 1.08 x 10^5

$\mathbf{\rho = \frac{108000}{8}}$

ρ = 13 500 kgm^{-3}

(d) What is the value of the atmospheric pressure of the surrounding? (1 mark)

Pressure at the top, P_t = atmospheric pressure, P_a = 8.0 x 10^4 Pa

Example 19

Figure 10.22 shows an insulated cylinder fitted with a pressure gauge, a heating coil and a frictionless piston of cross-sectional area 100 cm^2.

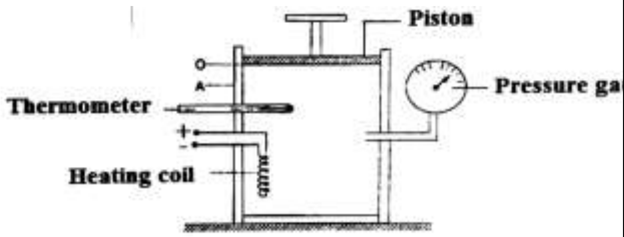


Figure 10.22

(a) While the piston is at position **O**, the pressure of the enclosed gas is 10 Ncm^{-2} at a temperature of 27 ^{0}C. 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 marks)

$\mathbf{P = \frac{F}{A}}$

$\mathbf{P = \frac{(10 \times 10)N}{100\ cm^2}}$

$\mathbf{P = 1\ Ncm^{-2}}$

Total pressure, $\mathbf{P_t = (10 + 1)\ Ncm^{-2}}$

$\underline{\mathbf{P_t = 11\ Ncm^{-2}}}$

(ii) State with a reason how the value obtained in (a) compares with the initial pressure. (2 marks)

The pressure increases since the volume of the gas increases.

(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 mark)

10 Ncm^{-2}

(ii) Determine the temperature of the gas in ^{0}C. (4 marks)

(Take g = 10 Nkg^{-1})

$\mathbf{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$ **(pressure law)**

$\mathbf{\frac{10}{300} = \frac{11}{T_2}}$

$\mathbf{T_2 = \frac{11 \times 300}{10}}$

$\mathbf{T_2 = 330\ K}$

$\mathbf{T_2 = (330 - 273)\ ^0C}$

$\underline{\mathbf{T_2 = 57\ ^0C}}$

Equation of State

- This is a general gas law relating the changes in pressure, volume and the absolute temperature as follows:

$$\frac{PV}{T} = \textbf{constant}, k$$

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

- The constant *k* depends on the:

(i) Type of the gas.

(ii) Quantity of the gas.

- When the amount of the gas is 1 mole, the equation changes to:

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

- **R** is a constant for all gases and is called the **universal gas constant.**

Example 20

A mass of 1 200 cm^3 of oxygen at 27 ^{0}C and a pressure of 1.2 atmospheres is compressed until its volume is 600 cm^3 and its pressure 3.0 atmospheres. Determine the temperature of the gas after compression in ^{0}C.

Solution

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \leftrightarrow T_2 = \frac{P_2 \times V_2 \times T_1}{P_1 \times V_1}$$

$$\mathbf{T_2} = \frac{\mathbf{3.0 \times 600 \times 300}}{\mathbf{1.2 \times 1\,200}}$$

$\mathbf{T_2 = 375\ K}$

$\mathbf{T_2 = (375 - 273)\ ^0C}$

$\underline{\mathbf{T_2 = 102\ ^0C}}$

Example 21

125 cm^3 of gas is collected at a temperature of 15 ^0C and pressure of 755 mm of mercury. Calculate the volume of the gas at s.t.p.

Note:

0^0C and **760 mmHg** are called **standard temperature** and **pressure (s.t.p.)**

Solution

$$\frac{\mathbf{P_1V_1}}{\mathbf{T_1}} = \frac{\mathbf{P_2V_2}}{\mathbf{T_2}} \leftrightarrow \mathbf{V_2} = \frac{\mathbf{P_1 \times V_1 \times T_2}}{\mathbf{P_2 \times T_1}}$$

$$\mathbf{V_2} = \frac{\mathbf{755 \times 125 \times 273}}{\mathbf{760 \times 288}}$$

$\underline{\mathbf{V_2 = 117.7\ cm^3}}$

GAS LAWS AND THE KINETIC THEORY OF MATTER

1. ***Boyle's Law***
   - If the volume of 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, i.e. the pressure is doubled.
   - Consequently, halving the volume of the gas doubles the pressure of the gas.
2. ***Charles' Law***
   - When the temperature of a gas rise, kinetic energy of the molecules of the gas increases.
   - The particles move faster and the rate of collision with the walls of the container increases.
   - Since the pressure is required to 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.
3. ***Pressure Law***
   - When the temperature of a gas rise, kinetic energy of the molecules of the gas increases.
   - Since volume is constant, the rate of collision with the walls of the container increases.
   - This increase in collisions increases the pressure of the gas.

Example 22

A house in which a cylinder containing cooking gas is unfortunately catches fire. The cylinder explodes. Use the kinetic theory of gases to explain the cause for the explosion. (2 marks)

High temperatures increase the kinetic energy of the gas molecules which leads to higher rate of collision with the cylinder walls, resulting in increase in pressure thereby causing explosion.

Example 23

When an inflated balloon is placed in a refrigerator it is noted that its volume reduces. Use the kinetic theory of gases to explain this observation. (3 mks)

Low temperature reduces the kinetic energy of the air molecules which leads to lower rate of collision with the walls of the balloon fabric which results to reduction of pressure.

Example 24

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 marks)

The hot temperature heats up the tyre which in turn heats up the air inside it. The air molecules gain more kinetic energy and move faster. Since the volume is constant, the molecules collide more frequently with the walls of the tyre which leads to greater change of momentum per unit time. This leads to an increase in pressure.

Limitations of Gas Laws

1. Kinetic theory of gases assumes that:

(a) **the size of the gas molecules is negligible.**

(b) the inter-molecular forces are negligible.

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

Note:

A gas that obeys the gas laws completely (or perfectly) is called **ideal** or **perfect gas**.