GAS LAWS

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Organizer

Objectives

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

(a) State Boyle's and Charles's laws.

(b) Describe experiments to illustrate Boyle's and Charles' laws.

(c) State and use the combined gas law to solve numerical problems.

(d) State Graham's law of diffusion and relate the rate of diffusion to relative molecular mass of a gas.

(e) Explain diffusion in terms of kinetic theory of matter.

GAS LAWS

Boyle's Law

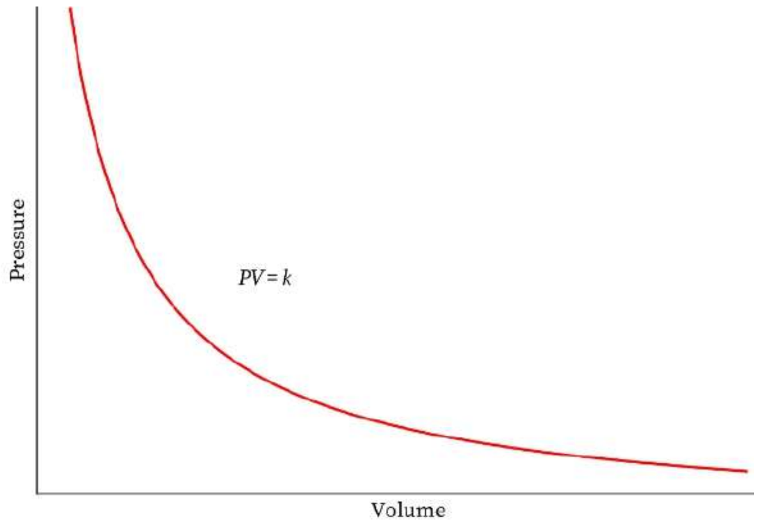
Boyle's law deals with the relationship between pressure and volume of a fixed mass of a gas when temperature is kept constant.

Pressure in a gas is as a result of the **collisions of the gas molecules with the walls of the container.** When the volume of the fixed mass of a gas is **decreased** through compression at constant temperature, the molecules travel a shorter distance to collide with the walls of the container, leading to increased number of collisions per unit time. The pressure of the gas therefore increases with the increased rate of collisions.

Boyle's law states that t*he volume of a given mass of a gas is inversely proportional to its pressure at constant temperature*.

What Boyle's law implies is that as the pressure increases, the volume decreases. The pressure of the gas inside the barrel of a pump is directly proportional to the physical pressure applied to compress the gas. A graph of the physical volume is a curve as shown below:.

The mathematical expression of Boyle's law is:



$$\text{Volume } \alpha \frac{1}{\text{Pressure}}$$

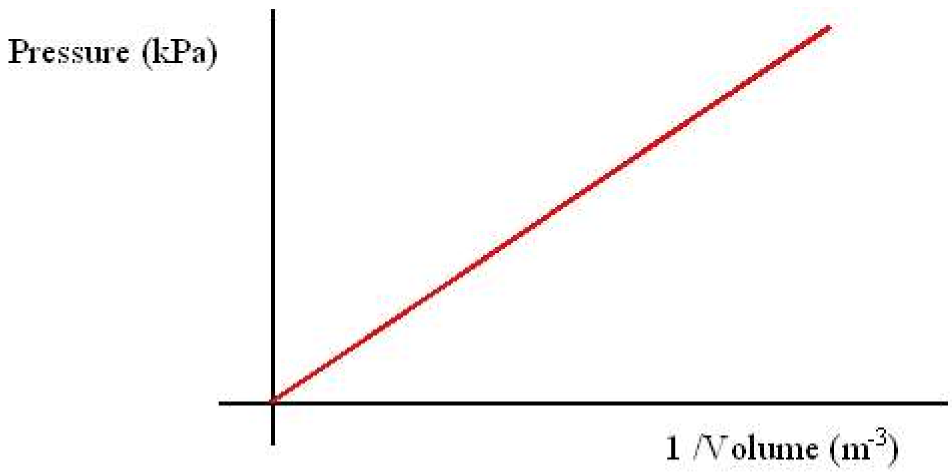
$$V \alpha = \frac{1}{P} \text{ or } P \alpha = \frac{1}{P}$$

Hence, VP = Constant

The expression implies that when the volume of a fixed mass of a gas changes from V_1 to V_2 its pressure also changes from P_1 to P_2. This leads to the general expression:

$$\mathbf{P_1V_1 = P_2V_2}$$

When a graph of pressure of a fixed mass of gas is plotted against the reciprocal of volume, a straight line is obtained.



The SI unit of pressure is the Pascal (Pa). It is equal to one Newton per square metre (NM^{-2}). Other units used to express pressure are atmospheres. One atmospheric pressure is equal to 760 mmHg pressure or 1.01325 × 10^5 Pascals. The SI unit of volume is cubic metres (m^3). One cubic metre is equal to 1.0 × 10^6 cubic centimetres (cm^3).

Worked Examples

1. A volume of 375 cm^3 of a gas has a pressure of 20 atmospheres. What will be its volume if pressure is reduced to 15 atmospheres?

Solution

From Boyle's law, $P_1 V_1 = P_2 V_2$

P_1 = 20 atmospheres, P_2 = 15 atmospheres, V_1 = 375 cm^3, V_2 = ?

Substituting for P_1, V_1 and P_2 the equation becomes;

$$20 \times 375 = 15 \times V_2$$

$$V_2 = \frac{20 \times 375}{15}$$

$$V_2 = 500 \text{ cm}^3$$

2. A given mass of gas occupies a volume of 200 cm^3 at a pressure of 5 atmospheres. At what pressure will the gas have a volume of 800 cm^3?

Solution

From Boyle's law, $P_1V_1 = P_2V_2$

P_1 = 5 atmospheres, P_2 = ?, V_1 = 200 cm^3, V_2 = 800 cm^3

$5 \times 200 = P_2 \times 800$

$$P_2 = \frac{5 \times 200}{800}$$

P_2 = 1.25 atmospheres

3. A certain mass of gas occupies 250 cm^3 at 25°C and 750 mmHg. Calculate its volume at 25°C if pressure changes to 760 mmHg in SI Units.

Solution

$V_1 = 250\ cm^3$, $V_1 = ?$, $P_1 = 750$ mmHg, $P_2 = 760$ mmHg

$$P_1V_1 = P_2V_2 \Rightarrow V_2 = \frac{P_1V_1}{2}$$

$$\therefore V_2 = \frac{750 \times 250}{760}$$

$$V_2 = 246.7 \times 10^{-6} m^3 = 2.467 \times 10^{-4} m^3$$

4. At a Constant temperature, a gas at 540 mmHg pressure occupies a volume of 300 litres. The gas is made to expand and occupy a volume of 600 litres. What is the new gas pressure in SI Units?

Solution

$P_1 = 540$ mmHg, $V_1 = 300\ cm^3$, $V_2 = 600\ cm^3$

$540 \times 300 = P_2 \times 600$

$$P_2 = \frac{540 \times 300}{600}$$

$$= 270 \text{ mmHg}$$

$$760 \text{ mmHg} = 1.01325 \times 10^5 \text{ Pascals}$$

$$270 \text{ mmHg} = \frac{1.01325 \times 10^5 \times 270P_a}{760}$$

$$= 3.6 \times 10^4 \text{ Pascals}$$

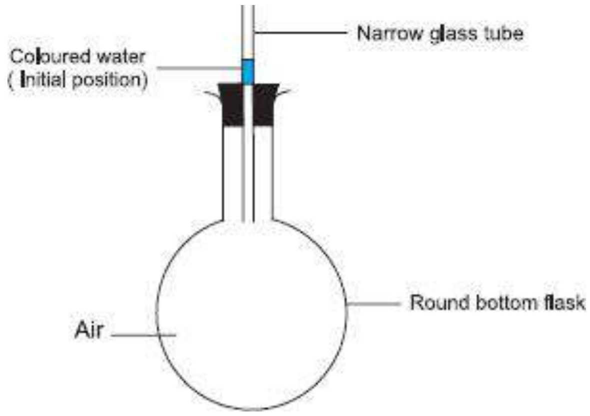
Charles' Law

Charles' Law deals with the relationship between the volume of a fixed mass of a gas with its temperature at constant pressure.

Charles's Law states that: *The volume of a given mass of a gas is directly proportional to its absolute temperature, its pressure being kept constant*.

How does the volume of a fixed mass of gas vary with temperature at constant pressure?

Fit a narrow glass tube into a rubber bung. Loosely fix the bung in a round bottomed flask. Introduce a drop of coloured water into the glass tube. When the drop is half way down the glass tube, firmly stopper the flask making it airtight. Note the position of the coloured water column in the tube. Immerse the flask in a trough of warm water. Observe and record what happens to the water column. Repeat the experiment using ice cold water in a trough. Record your observations.



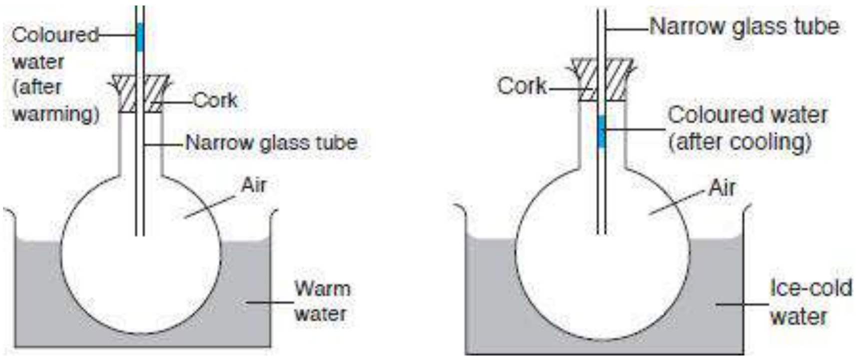
Volume of a given mass of a gas

Questions

1. **What happens to the water column when the flask is warmed or cooled?**

When the flask is immersed in a trough of warm water, the column of coloured water first drops then moves up steadily.

When the flask is immersed in warm water the coloured water column in the capillary tube drops initially because the flask expands before the gas inside absorbs heat to start expanding.



2. How does the change in temperature affect the volume of the fixed mass of gas?

   Charles's Law: *The volume of a given mass of a gas is directly proportional to its absolute temperature, its pressure being kept constant*.

3. Which factor is kept constant in this experiment?

   Pressure is kept constant.

4. Explain the observations made during the experiment in terms of the kinetic theory.

When the gas inside heats up, it expands and pushes up the water column in the capillary tube. When a fixed mass of a gas is heated at constant pressure, its volume increases to counter balance the constant pressure. The heat energy increases the kinetic energy of the gas molecules. This leads to increased rate of collisions with the walls of the container causing an increase in the gas pressure.

Mathematically, Charles's Law is expressed as; Volume α absolute temperature. V α T

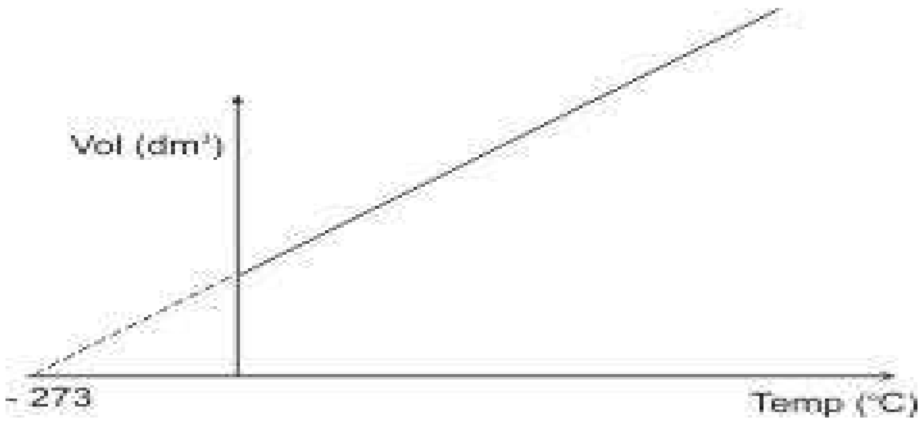
(Where V is the volume of the gas and T its absolute temperature in Kelvin).

$$\text{Thus, } \frac{V}{T} = \text{constant}$$

When the volume of a fixed mass of a gas changes from V_1 to V_2, its absolute temperature changes from T_1 to T_2 leading to the expression:

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

A graph of volume of a fixed mass of a gas against temperature is a straight line. The graph intercepts the temperature axis at −273 °C when extrapolated as shown below.



A gas **whose volume is theoretically zero at -273°C** is referred to as an **ideal gas.**

In reality, a gas **would have turned into a liquid** before reaching a temperature of **−273°C.**

The **temperature at which the volume of a gas is assumed to be zero** is called **absolute zero.**

From the absolute zero temperature, the absolute temperature scale is derived. The lowest temperature on the absolute scale is **zero Kelvin (0 K). The Kelvin (K) is the SI unit of the temperature on the absolute scale.** The equivalence of **−273°C** on the Kelvin scale is **0 K**.

Temperature inter conversion from one scale to the other is done as follows:

1 **To convert temperature in degrees Celsius (°C) to Kelvin (K), add 273. For example, to convert 204 °C to Kelvin add 273 this;**

T = 204 + 273

= 477K

2. **To convert temperature in the Kelvin scale to degrees Celsius, subtract 273.**

For example, convert 405 K to degrees celsius (°C)

t = 450 – 273

= 177°C.

Worked Examples

1. Convert the temperatures below to the absolute scale:
   (i) 0°C
   (ii) 25°C

(iii) –20°C

Solution

T = t + 273

(i) 0 + 273 = 273 K

(ii) 25 + 273 = 298 K

(iii) –20 + 273 = 253 K

Note:

Temperature on the Kelvin scale is denoted by T, while on the celsius scale it is denoted by t.

2. Convert the temperature recorded below in Kelvin to temperature in degrees celsius (°C):

(i) 0 K

(ii) 250 K

(iii) 273 K

Solution

t = T – 273

t = T – 273

t = T – 273

(i) 0 – 273 = – 273°C

(ii) 250 – 273 = – 23°C

(iii) 273 – 273 = 0°C

3. A gas occupies 450 cm^3 at 27°C. What volume would the gas occupy at 177°C, if its pressure remains constant?

Solution

From Charles's Law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Converting the temperature to absolute temperature then substituting for V_1, T_1 and T_2 then:

$$\frac{V_2}{450} = \frac{450}{300}$$

$$V_2 = \frac{450 \times 450}{300} = 675 \text{ cm}^3$$

4. At a temperature of 57°C, nitrogen gas occupies a volume of 750 cm^3. At what temperature will the gas occupy 100 cm^3. Express the answer in degrees celsius.

Solution

From Charles's Law: $\frac{V_2}{T_2} = \frac{V_1}{T_1}$

$$\frac{100}{T_2} = \frac{750}{330}$$

$$T_2 = \frac{100 \times 330}{750} = 44\ K$$

Converting to °C celsius: $t = T - 273$

$$= 44 - 273$$

$$= 229°C$$

3.Combined Gas Law

The combined gas law deals with the variation in the volume of a fixed mass of a gas with respect to changes in temperature and pressure.

The mathematical expression of Charles's Law, V α T can be combined with that of Boyle's law V α $\frac{1}{P}$ to obtain the expression:

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

Therefore, PV α T

Hence, $\frac{PV}{T}$ = Constant

If a fixed mass of a gas of volume V_1 exerts a pressure P_1 at absolute temperature T_1, the expression may be written as:

$$\frac{P_2V_2}{T_2} = \text{Constant}$$

Suppose the same mass of gas has a volume, $V_{2,}$ and exerts a pressure, $P_{2,}$ at absolute temperature, $T_{2,}$ then the expression becomes:

$$\frac{P_1V_1}{T_1} = \text{Constant}$$

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

This is the **ideal gas equation.** It enables the volume of a gas to be obtained under any conditions of temperature and pressure provided its volume under some other conditions of temperature and pressure is known.

Standard Conditions

There are two conditions considered when comparing volumes of gases:

1. **Standard temperature and pressure (s.t.p).** The s.t.p conditions refer to a temperature of 273 K and a pressure of 760 mmHg (1 atmosphere).

2. **Room temperature and pressure (r.t.p).** This refers to a temperature of 298 K and a pressure of 760 mmHg.

It should be seen that from the standard conditions, all computations of temperature should be expressed in Kelvin.

Worked Examples

1. **What will be the volume of a given mass of oxygen at 25°C if it occupies 100 cm^3 at 15°C? (pressure remains constant).**

Solution

Let the initial volume, temperature and pressure be V_1, T_1, P_1 respectively, and the final be V_2, T_2, P_2. From the gas equation;

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

Where, $V_1 = 100$ cm^3, $T_1 = 15 + 273 = 288$K and $P_1 = P_2 =$ constant.

$$V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2} \text{ since } P_1 = P_2$$

$$\text{Then, } V_2 = \frac{V_1T_2}{T_1} = \frac{100 \times 298}{288} = 103.5 \text{ cm}^3$$

2. **A given mass of a gas occupies 20 cm^3 at 25°C and 670 mmHg pressure. Find out the volume it will occupy at:**
   (a) 10°C and 335 mmHg
   (b) 0°C and 760 mmHg

Solution

(a) From the gas equation,

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

Where, $P_1 = 670$ mmHg, $V_1 = 20$ cm^3,

$T_1 = 25 + 273 = 298$ K, $T_2 = 10 + 273 = 283$ K, $P_2 = 335$ mmHg, $V_2 = ?$

$$\text{Therefore, } \frac{670 \times 20}{298} = \frac{335 \times V_2}{283} \qquad V_2 = \frac{670 \times 20 \times 283}{288 \times 335} = 38 \text{ cm}^3$$

(b) $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Where, $P_1 = 670$ mmHg, $V_1 = 20$ cm^3, $T_2 = 273$ K, $V_2 = ?$

$$\text{Therefore, } \frac{670 \times 20}{298} = \frac{760 \times V_2}{273} \qquad V_2 = \frac{670 \times 20 \times 273}{298 \times 760} = 16 \text{ cm}^3$$

Some Applications of Gas Laws

The effects of changes in pressure, volume and temperature on a fixed mass of a gas have been used in a wide range of applications which include:

1. Inflating tyres, balls and balloons appropriately depending on the prevailing temperature conditions.
2. Designing of aerosol cans and tear-gas canisters which contain a gas compressed under pressure to act as a propellant of liquid contents in the cans.
3. Regulation of pressure in an aircraft for comfortable in flight environment at high altitude.

Diffusion and the Graham's Law

Diffusion is the process by which **particles spread out from a region of high concentration to regions of low concentration.**

The scent of a strong perfume reaches all corners of a room as soon as the container is opened because of diffusion. The **inter–molecular forces of attraction in gases are very weak due to their large inter-molecular distances. A gas therefore, always spreads out to fill up all the space available. The perfume spreads from an area of high concentration to areas where its concentration is low.**

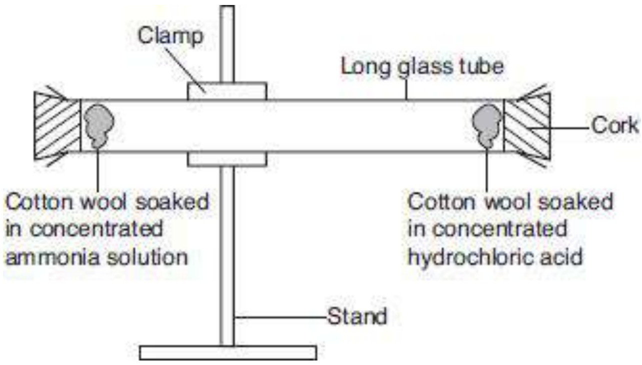
The spreading out of gas particles in air takes a shorter time than solid particles in a solvent. This is because gas particles are far apart and have more kinetic energy than the liquid particles.

Graham's law of diffusion.

Different gases have different rates of diffusion.

According to **Graham's Law**, **the rate of diffusion of a gas is inversely proportional to the square root of its density *under* the same conditions of temperature and pressure**

To verify this, two gases, ammonia and hydrogen chloride can be allowed to diffuse inside a combustion tube as shown below.



Questions

1. **What observations are made in the glass tube and after how long?**

   A white solid is formed in the tube closer to the end where the cotton wool soaked in concentrated hydrochloric acid was placed after about 5 minutes.
2. **Which gas covered a longer distance?**

   Ammonia
3. **Explain the observations made in the glass tube.**

Concentrated ammonia solution generates ammonia gas while concentrated hydrochloric acid generates hydrogen chloride gas. Ammonia and hydrogen chloride gases diffuse in air in the tube, and when their molecules meet, they react to form white, solid ammonium chloride.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s$

4. **Determine the molecular masses of ammonia (NH_3) and hydrogen chloride (HCl). N = 14.0, H = 1.0, Cl = 35.5.**

   Ammonia =14+3=17g/mol

   Hydrogen chloride =1+35.5=36.5g/mol

5. **Calculate the rate of diffusion of:**

   (a) Ammonia gas.

The distance covered by ammonia was 12 cm. The rate of diffusion of ammonia gas in air is $\frac{12\text{ cm}}{5\text{ minutes}} = 2.4$ **cm/minute**

   (b) Hydrogen chloride gas.

The distance covered by hydrogen chloride was 8 cm within the same time interval of 5 minutes.

The rate of diffusion of hydrogen chloride gas in air is

$\frac{8\text{ cm}}{5\text{ minutes}} = 1.6$ **cm/minute.**

6. **What is the relative rate of diffusion of ammonia to hydrogen chloride gas in air?**

The relative rate of diffusion of ammonia gas compared to hydrogen chloride gas in air is:

$$\frac{\text{Rate of diffusion of NH}_3}{\text{Rate of diffusion of HCl}} = \frac{2.4\text{ cm/minutes}}{1.6\text{ cm/minuts}} = 1.5$$

This means ammonia diffuses 1.5 times faster than hydrogen chloride. This is because ammonia gas is less dense than hydrogen chloride gas. Therefore, gases with low densities diffuse faster than those with high densities.

The mathematical expression of Graham's Law is:

$$\text{Rate (R)}\ \alpha = \frac{1}{\sqrt{\text{density(d)}}}$$

Meaning that:

$$\text{Rate (R)} = \frac{\text{Constant}}{\sqrt{\text{density(d)}}}$$

$$R = \frac{\text{Constant}}{\sqrt{d_A}}$$

When the rates of diffusion of two gases A and B are compared, the equations are:

(i) Rate of diffusion of gas A = $\frac{\text{Constant}}{\sqrt{\text{density of gas A}}}$

$$R_A = \frac{\text{Constant}}{\sqrt{d_A}}$$

(ii) Rate of diffusion of gas B = $\frac{\text{Constant}}{\sqrt{\text{density of gas B}}}$

$$R_B = \frac{\text{Constant}}{\sqrt{d_B}}$$

In equation (i), $R_A\sqrt{d_A} = \text{Constant}$

In equation (ii), $R_B\sqrt{d_B} = \text{Constant}$

Therefore, $R_A\sqrt{d_A} = R_B\sqrt{d_B}$

$$\frac{R_A}{R_B} = \sqrt{\frac{d_B}{d_A}}$$

Since density is directly proportional to molecular mass, Graham's law can also be expressed as:

$$\text{Rate} = \frac{\text{Constant}}{\sqrt{\text{molecular mass}}}$$

Therefore, if the rate of diffusion of two gases A and B are compared, then:

$$\frac{R_A}{R_B} = \sqrt{\frac{M_B}{m_A}}$$

Where M_A and M_B are relative molecular masses of the gases A and B respectively. Since Rate is inversely proportional to time, (Rate α $\frac{1}{\text{Time}}$).

It means that it is also possible to compare the time taken for **equal volumes** of two gases to diffuse under similar conditions:

$$R_A = \frac{\text{Constant}}{\text{Time of diffusion of gas A}} \text{ and } R_B = \frac{\text{Constant}}{\text{Time of diffusion of gas B}}$$

This means that:

R_AT_A = Constant, and R_BT_B = Constant

(R = rate, T = time)

Therefore,

$R_AT_A = R_BT_B$ and $\frac{R_A}{R_B} = \frac{T_B}{T_A}$

Where T_A and T_B are the times of diffusion of gases A and B respectively.

$$\text{But } \frac{R_A}{R_B} = \sqrt{\frac{d_B}{d_A}} = \sqrt{\frac{M_B}{M_A}}$$

$$\text{Therefore, } \frac{T_B}{T_A} = \sqrt{\frac{d_B}{d_A}} \text{ and } \frac{T_B}{T_A} = \sqrt{\frac{M_B}{M_A}} \qquad \frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$$

Worked Examples

1. **Equal volumes of carbon(II) oxide and carbon(IV) oxide are allowed to diffuse through the same medium. Calculate the relative rate of diffusion of carbon(II) oxide. (C = 12.0, Oxygen = 16.0)**

Solution

Relative molecular mass (M_r) of CO = 12 + 16 = 28

Relative molecular mass (M_r) of CO_2 = 12 + 32 = 44

$$\frac{R_{CO}}{R_{CO_2}} = \sqrt{\frac{M_r CO_2}{M_r CO}}$$

$$\frac{R_{co}}{R_{CO_2}} = \sqrt{\frac{44}{28}}$$

$$\frac{R_{co}}{R_{CO_2}} = 1.254$$

Carbon(II) oxide diffuses 1.254 times faster than carbon(IV) oxide.

2. **If it takes 20 seconds for 200 cm^3 of oxygen gas to diffuse across a porous plug. How long will it take an equal volume of sulphur(IV) oxide to diffuse across the same plug? (O = 16.0, S = 32.0).**

Solution

$$\frac{T_{O_2}}{T_{SO_2}} = \sqrt{\frac{M_r O_2}{M_r SO_2}} \Rightarrow \frac{20}{T_{SO_2}} = \sqrt{\frac{32}{64}}$$

$$\frac{20}{T_{SO_2}} = \frac{1}{\sqrt{2}}$$

$$T_{SO_2} = 20\sqrt{2} = 28.3 \text{ seconds}$$

3. **Determine the molecular mass of the gas Y which diffuses $1\frac{1}{2}$ times slower than Oxygen. 0 = 16.0**

$$\frac{R_Y}{R_{O_2}} = \sqrt{\frac{M_rO_2}{M_rY}}$$

$$\frac{1}{1.5} = \sqrt{\frac{32}{M_{r_Y}}}$$

$$\frac{1}{2.25} = \frac{32}{M_{r_Y}}$$

$$M_{r_Y} = 2.25 \times 32$$

$$M_{r_Y} = 72$$

4. **If it takes 30 seconds for 100 cm^3 of carbon(IV) oxide to diffuse across a porous plate. How long will it take 150 cm^3 of nitrogen(IV) oxide to diffuse across the same plate under similar conditions? (C = 12.0, N = 14.0, O = 16.0).**

Solution

$$\frac{R_{CO_2}}{R_{NO_2}} = \sqrt{\frac{M_rNo_2}{M_rCo_2}} \quad \text{But } RCO_2 = \frac{100\text{cm}^3}{30\text{s}} = 3.33\text{cm}^3 \text{ per second}$$

$$\frac{3.33}{R_{NO_2}} = \sqrt{\frac{46}{44}}$$

$$\frac{3.33}{R_{NO_2}} \; 1.0225$$

$$R_{NO_2} = \frac{3.33}{1.0225} = 3.26 \text{ cm}^3 \text{ per second}$$

Therefore, time taken for NO_2 to diffuse is:

$$\frac{150 \text{ cm}^3}{3.26 \text{ cm}^3 \text{sec}^{-1}} = 46 \text{ seconds.}$$

Alternative Working

Since equal volumes must be compared:
100 cm^3 of CO_2 takes 30 seconds

$$\therefore 150 \text{ cm}^3 \text{ of } CO_2 \text{ will take } \frac{30}{100} \times 150 = 45 \text{ seconds}$$

$$\frac{T_{CO_2}}{T_{NO_2}} = \sqrt{\frac{M_rCO_2}{M_rNO_2}}$$

$$\frac{45}{T_{NO_2}} = \sqrt{\frac{44}{28}} \Rightarrow \frac{45}{T_{NO_2}} = 0.978$$

$$T_{NO_2} = \frac{45}{0.978} = 46 \text{ seconds}$$

5. **Calculate the relative rate of diffusion of ammonia gas compared to that hydrogen chloride gas under the same conditions of temperature and pressure.**

From the expression:

$$\frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$$

$$\text{Then; } \frac{R_{NH_3}}{R_{HCl}} = \sqrt{\frac{35.5}{17}} = 1.465$$

This ratio implies that ammonia diffuses 1.5 times faster than hydrogen chloride gas.

Review Questions

1. 2006 Q 3 P1

60cm^3 of oxygen gas diffused through a porous partition in 50 seconds. How long would it take 60cm^3 of sulphur (IV) oxide gas to diffuse through the same partition under the same conditions? (S= 32.0, 0 = 16.0)
(3marks)

2. 2007 Q 12 P1

(a) State the Charles law (1mark)

(b) The volume of a sample of nitrogen gas at a temperature of 291 K and 1.0x10^5 Pascals was 3.5 x 10^{-2}m^3. Calculate the temperature at which the volume of the gas would be 2.8 x 10^{-2}m^3 at 1.0 x 10^5 Pascals.
(2marks)

3. 2008 Q 1 P1

A small crystal of potassium manganate (VII) was placed in a beaker water. The beaker was left standing for two days without shaking. State and explain the observations that were made.
(2marks)

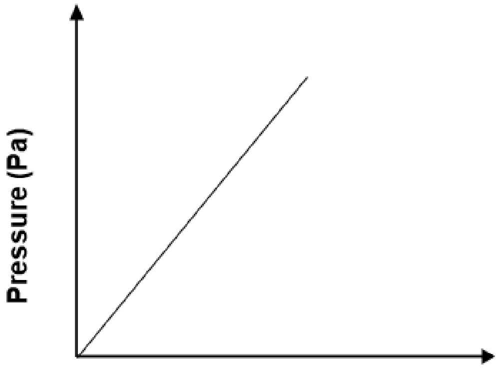
4. 2008 Q 20 P1

(a) State the Graham's law diffusion. (1mark)

(b) The molar masses of gases **W** and **X** are 16.0 and 44.0 respectively. If the rate of diffusion of **W** through a porous material is 12cm^{3}s^{-1}, calculate the rate of diffusion of **X** through the same material.
(2marks)

5. 2009 Q 26 P1

The graph below shows the relationship between pressure and the temperature of a gas in a fixed volume container.



(a) State the relationship between pressure and temperature that can be deduced from the graph. (1 mark)

(b) Using kinetic theory, explain the relationship shown in the graph. (2 marks)

6. 2010 Q8 P1

The pressure of nitrogen gas contained in a 1 dm^3 cylinder at -196 °C was 10^7 Pascals. Calculate the:

(a) Volume of the gas at 25 °C and 10^5 Pascals. (1½ marks)

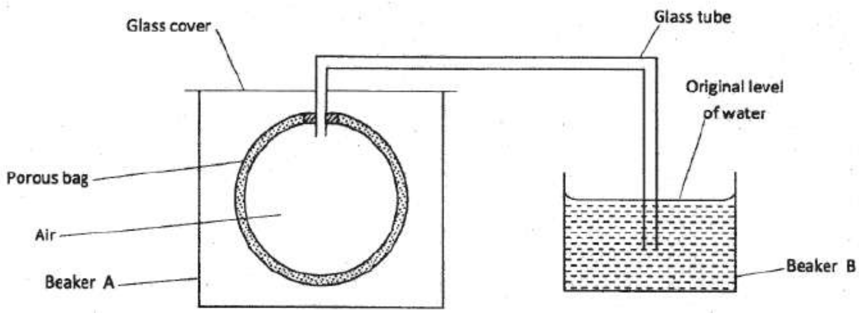
(b) Mass of nitrogen gas.
(Molar volume of gas is 24 dm^3, N = 14.0) (1 ½ marks)

7. 2011 Q6 P1

A certain mass of gas occupies 0.15 dm^3 at 293K and 98,648.5 Pa. Calculate its volume at 101,325 Pa and 273K. (2 marks)

8. 2012 Q19 P1

The set up shown below was used to investigate a property of hydrogen gas.



State and explain the observation that would be made in the glass tube if beaker **A** was filled with hydrogen gas. (3 marks)

9. 2013 Q14 P1

(a) State the Charles' law. (1 mark)

(b) A certain mass of gas occupies 146 dm^3 at 291 K and 98.31 kPa. What will be its temperature if its volume is reduced to 133 dm^3 at 101.325 Pa?

(2 marks)

10. 2014 Q6 P1
100cm^3 of a sample of ethane gas diffuses through a porous pot in 100 seconds. What is the molecular mass of gas Q if 1000 cm^3 of the gas diffuses through the same porous pot in 121 seconds under the same conditions?
(C-12.0, H=1.0) (3 marks)

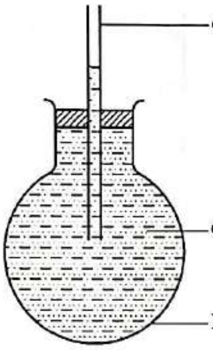
11. 2015 Q4 P1
(a) State the Boyle's Law. (1 mark)
(b) A gas occupies 500cm^3 at 27 °C and 100,000 Pa. What will be its volume at 0 °C and 101,325 Pa? (2 marks)

12. 2016 Q20 P1
60cm^3 of oxygen gas diffused through a porous partition in 50 seconds. How long would it take 60cm^3 sulphur (IV) oxide gas to diffuse through the same partition under the same condition (S=32.0, O=16.0) (3 marks)

13. 2017 P1 Q6.
(a) State Charles' Law. (1 mark)
(b) Explain why the pressure of a fixed mass of a gas increases, when the volume of the gas is reduced at constant temperature. (2 marks)

14. 2018 P1 Q 6.
(a) State Graham's law of diffusion. (1 mark)
(b) Explain why a balloon filled with helium gas deflates faster than a balloon of the same size filled with argon gas. (2 marks)

15. 2018 P2 Q3(b)
Use the set-up in **Figure 3** to answer the questions that follow. The flask was covered with a cloth that had been soaked in ice-cold water.



(i) State the observation made on the coloured water. Explain. (2 marks)
(ii) Name the gas law illustrated in **Figure 3.** (1 mark)

16. 2019 P1 Q26.
140cm^3 of nitrogen gas diffuses through a membrane in 70 seconds. How long will it take 200 cm^3 of carbon (IV) oxide gas to diffuse through the same membrane under the same conditions of temperature and pressure. (3 marks)