

Dalton's law of partial pressure

It states that "the total pressure exerted by a mixture of gases that do not chemically react is equal to the sum of the partial pressures of gas components at constant temperature."

For a mixture of two gases A and B in the same vessel, the total pressure, P_T in the vessel is given by the expression;

$$P_T = P_A + P_B$$

Where P_A and P_B are the partial pressures of gases A and B respectively.

Partial pressure

Is the pressure that a gas would exert if it alone occupied the whole volume that is actually occupied by the mixture of gases of which it is a part.

Or

Is the pressure the gas alone would exert if it occupied the volume that was initially occupied by the mixture of the gases of which it is a part.

Partial pressure of a gas component is given by the product of the mole fraction of the gas and the total pressure of the gases. E.g. $P_A = X_A P_T$, where

P_A is the partial pressure of gas A

X_A is the mole fraction of gas A

P_T is the total pressure exerted by a mixture of gases

Mole fraction

It is the ratio of the number of moles of a particular gas to the total number of moles of gases in the mixture

Consider two gases A and B in the same vessel, then mole fractions A and B are obtained as follows;

$$X_A = \frac{n_A}{n_A + n_B} \quad \text{and} \quad X_B = \frac{n_B}{n_A + n_B}$$

Where X_A and X_B are mole fractions of gases A and B respectively

n_A and n_B are numbers of moles of gases A and B respectively

$n_A + n_B$ is the total number of moles of gases A and B in the mixture

Thus, partial pressures, P_A and P_B , of gases A and B are obtained as below;

$$P_A = \frac{n_A}{n_A + n_B} \times P_T \quad \text{and} \quad P_B = \frac{n_B}{n_A + n_B} \times P_T$$

Also, $X_A + X_B = 1$ therefore, $X_A = 1 - X_B$ and $X_B = 1 - X_A$

$$\Rightarrow P_A = (1 - X_B)P_T \quad \text{and} \quad P_B = (1 - X_A)P_T$$

Questions:

- 2.6g of carbon dioxide and 4.8g of oxygen were mixed in a one litre container. If the total pressure is 100atm. Calculate the partial pressure of each gas in the container.
- A mixture of hexane (C_6H_{14}) and Heptane (C_7H_{16}) exerts a pressure of 95kPa. If the partial pressure of Hexane is 50kPa, determine the mole fraction of Heptane in the gas mixture.
- 4dm³ of O₂ at a pressure of 200KPa and 1 dm³ of N₂ at a pressure of 200KPa are introduced in a 2 dm³ vessels. What is the total pressure in the vessel?
- Calculate the pressure that is exerted by 2 moles of hydrogen in a volume of 20 litres at 20°C
- A mixture of gases at s.t.p. contains 65% nitrogen, 15% carbon dioxide and 20% oxygen. What is the partial pressure of each gas in kPa (standard pressure = 1 atm = 101.325 kPa)
- 267cm⁻³ of a gas at 18°C and 100400 Pa was found to weigh 0.162g. Calculate the relative molecular mass of the gas.
- The partial pressures of components of a mixture of gases are 26.64kPa for O₂, 34kPa for N₂ and 42.6kPa of H₂. Find the percentages of O₂ by volume in the mixture.

Kinetic theory of gases

It states that gases consist of molecules that are in constant random motion in all directions. The following are the assumptions of kinetic theory of gases.

1. Gases are made up of tiny particles called molecules.
2. Molecules are in state of rapid random motion colliding with each other and with the walls of the container hence constituting pressure.
3. Gas molecules undergo perfectly elastic collisions.
4. The average kinetic energy of gas molecules is directly proportional to absolute temperature and independent of the chemical nature of the gas.
5. The forces of attraction between the gas molecules and walls of the container are negligible.

Real gases/ Non-ideal gases.

Real gases are gases that do not obey Boyle's law and ideal gas equation.

Assumption made for non-ideal gases.

- The forces of attraction between the gas molecules are pronounced.
- Gas molecules occupy an appreciable volume compared to volume of the container.
- Gas molecules have low kinetic energy
- Gas molecules may not execute elastic collision
- Pressure exerted by the real gas molecules may not be accounted for by the mean collision between them and walls of the container.

The deviation from ideal behavior increases as the molecular sizes (masses) increase and also they deviate at **High pressure** (when compressed) and **low temperatures**.

Factors responsible for the behaviour of non-ideal gases are;

- 1) **Intermolecular forces of attraction** (At low temperatures)

At **low temperatures**, the Van der Waal's forces of attraction become more pronounced because the gas molecules have insufficient kinetic energy to overcome the forces of attraction. The molecules come closer together and the velocity with which they strike the walls of the container reduces, thus, reducing pressure below that of an ideal gas hence resulting in deviation from ideal gas equation. i.e. $PV < nRT$. Thus, the effect of **molecular attraction** is greater than **molecular volume**.

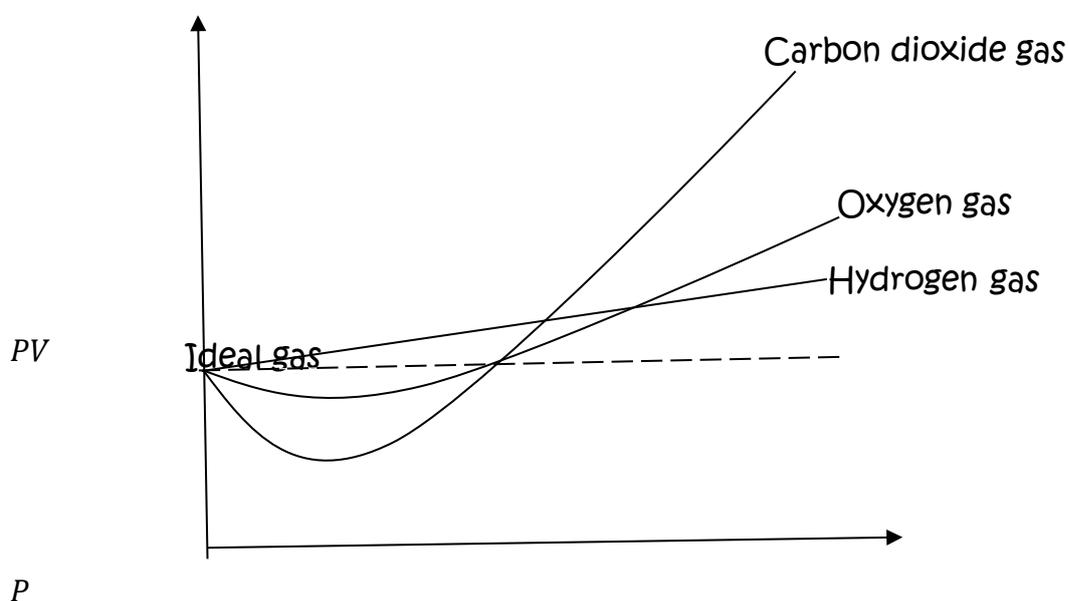
- 2) **Molecular volume** (At high pressures)

At high pressure, the molecules of a gas are very close together and the volume the gas molecules occupy is appreciable (reasonable) and cannot be neglected compared to the volume of the container. Thus, $PV > nRT$ for n moles of the gas and the effect of molecular volume is greater than molecular attraction.

The deviation can be shown in two ways;

a) By plotting a graph of PV against P .

Note: For an ideal gas, PV is constant at all conditions.



From the graph, deviation from ideal gas behavior increases as pressure increases

At low pressure, the gases tend to behave ideally and at such a low pressure the molecules are very far apart so that the intermolecular forces of attraction are not exerted in them.

The volume of the container is also very large such that the volume occupied by the gaseous molecules is negligible.

Hydrogen tends to approximate ideality because of its low molecular size (mass).

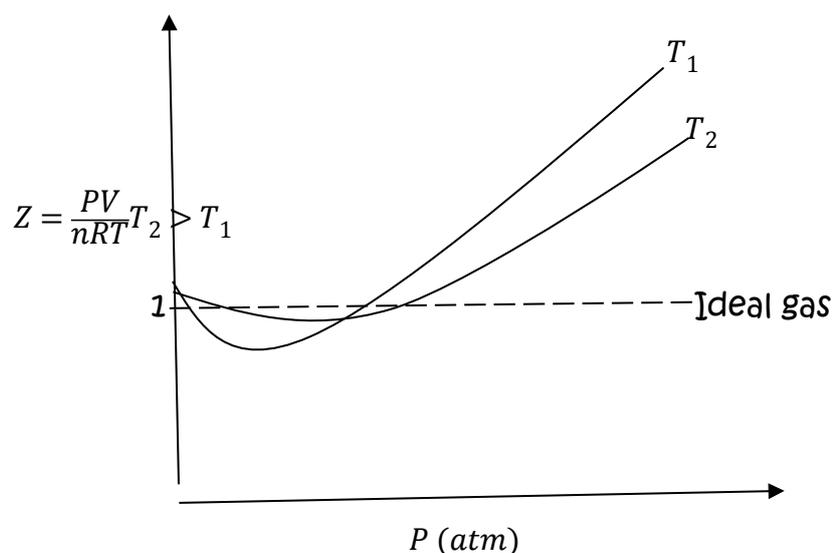
Carbon dioxide shows greater deviation from ideal behavior than even oxygen because the carbon dioxide molecules are bigger with stronger

intermolecular forces of attraction (Van der Waal's forces) compared to the molecules of the other gases.

At low pressures pressure, the curves for oxygen and carbon dioxide fall because the effect due to molecular attraction is greater than that due to molecular volume causing a reduction in pressure on the walls of container hence the value of P in the expression of PV reduces causing a fall in curve for oxygen and carbon dioxide.

As pressure increases (at very high pressures), the curves for oxygen and carbon dioxide raise because the effect due to molecular volume is greater than that due to intermolecular force of attraction (molecular attraction) making the volume of gas molecules more significant compared to the volume of the container hence an increase in the value of PV .

A graph of $\frac{PV}{nRT}$ against P for the same gas at different temperatures



The ratio $Z = \frac{PV}{nRT}$ is known as the **compressibility factor** and its value is 1 for ideal gases.

Compressibility factor is the measure of deviation of a gas from ideal gas behaviour. At low pressure, there is a slight fall in the value of Z because the effect of molecular attraction outweighs the deviation due to molecular volume. The value of Z increases at higher pressures as the effect of molecular volume outweighs that of molecular attraction.

Correction factor for real gases (The Van der Waal's equation)

A real gas occupies a significant volume in relation to the volume of a container. Consider a large vessel of volume, $V \text{ cm}^3$. If the volume occupied by gas molecules is $b \text{ cm}^3$, the volume where other molecules can move or the compressible part is $(V - b) \text{ cm}^3$. If there are n number of molecule of the real gas in the container, then the volume becomes $(V - nb) \text{ cm}^3$. Therefore, $PV = nRT$ also becomes $P(V - nb) = nRT$.

For a real gas, the gas molecules attract each other and a molecule in the interior is equally attracted at all sides. The velocity with which the molecules strike the wall of the container reduces thus reducing pressure below that of an ideal gas.

Considering two molecules, one in the interior of the container and the other around the wall, "the pressure reduction is proportion to the product of concentration of molecules in the interior of the container and concentration of molecules around the wall."

$$\text{Pressure reduction} \propto \frac{n}{v} \times \frac{n}{v} \quad \Rightarrow \text{Pressure reduction} = \frac{an^2}{v^2}$$

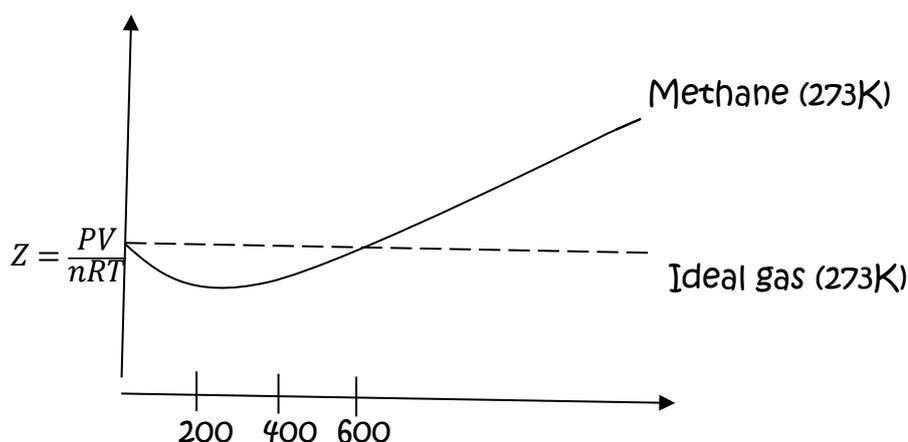
Since the observed/actual pressure is smaller than that expected for an ideal gas by $\frac{an^2}{v^2}$, Vander Waal added the term $\frac{an^2}{v^2}$ to the observed pressure to correct for pressure and P in the ideal gas becomes $(P + \frac{an^2}{v^2})$.

Thus, the expression $P(V - nb) = nRT$ becomes $(P + \frac{an^2}{v^2})(v - nb) = nRT$.

Therefore, $(P + \frac{an^2}{v^2})(v - nb) = nRT$ is the Van der Waal's Equation

a and b are constants indicating correction factors for volume and pressure respectively.

Question: A graph of $\frac{PV}{nRT}$ against pressure for 1 mole of methane is shown below.



- Give the reason for the behaviour of methane gas at 273K when the pressure is lower than 200 atm.
- Give a reason for the behaviour of methane gas at 273K when the pressure is higher than 200 atm.
- On the same axis, sketch a graph of ethane gas at 273K and give reasons for your sketch. Label the curve X.

Question: UNEB 2012 P/2 Qn:1

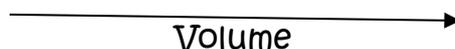
- By plotting a graph of pressure against volume at various temperatures.

A graph of pressure against volume at a given temperature is an **isothermal**.

An isothermal of an ideal gas will always take up the shape below while those of real gases will differ a certain temperature and below since they do not obey Boyle's law at all times. Real gases only obey Boyle's law at **high temperatures** and **low pressures**.



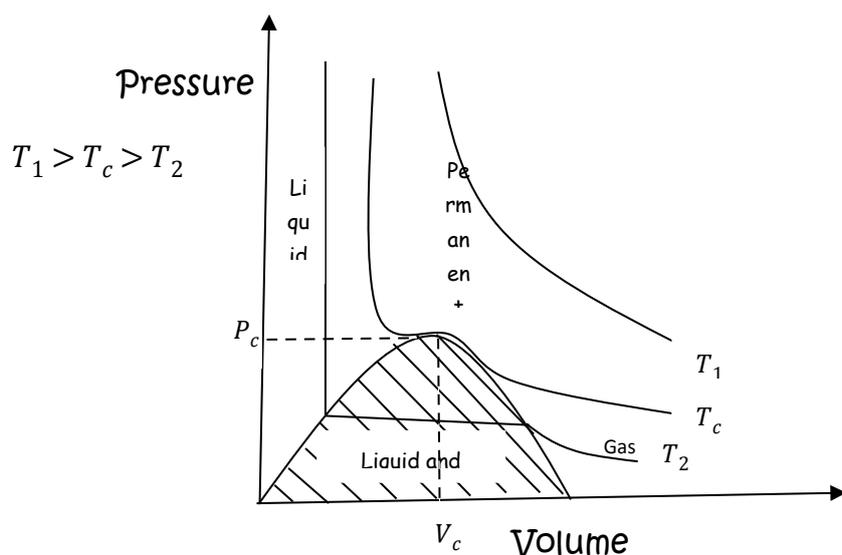
Pressure



LIQUEFACTION/ CONDENSATION OF GASES

When a gas is sufficiently compressed, it condenses into a liquid.

Conditions under which gases are liquefied were studied by Andrew who subjected Carbon dioxide to various pressures while changing temperature and he obtained the following isothermals.



When pressure is increased on a fixed mass of carbon dioxide gas at temperature T_1 , the volume decreases and the isothermal at this temperature obeys Boyle's law i.e. at T_1 the gas exhibits an ideal behaviour.

At temperature T_c , increase in pressure reduces volume of carbon dioxide gas but at point B, the gas begins to liquefy. Above T_c , no liquefaction takes place, T_c is referred to as the **critical temperature**.

Critical temperature is the maximum temperature at which a gas can be liquefied and beyond which no liquefaction takes place however much pressure is applied.

The minimum pressure which is just enough to liquefy one mole of a gas at its critical temperature is called **Critical pressure**.

Critical volume is the volume occupied by one mole of a gas at its critical temperature and pressure.

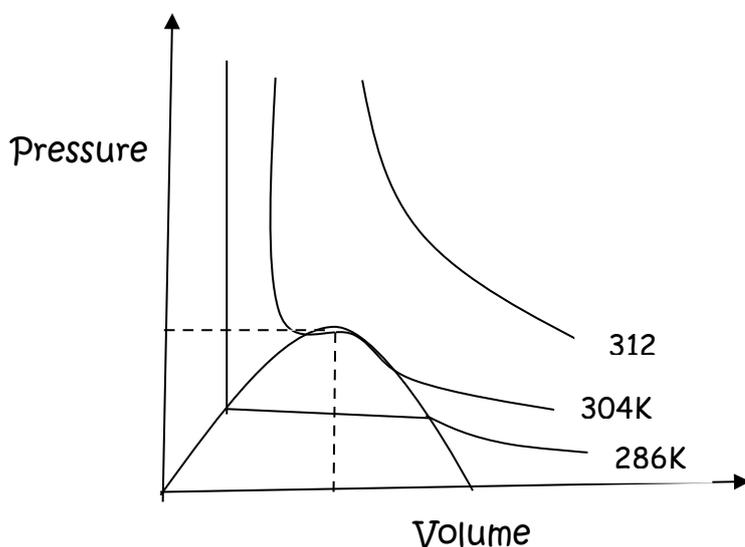
At point *W*, carbon dioxide is in gaseous state and as pressure is gradually increased, the volume decreases according to Boyle's law therefore between *W* and *X*, the gas behaves ideally.

At *X*, the gas begins to liquefy and there is a rapid decrease in volume at almost constant pressure. This is because the liquid formed has very high density.

At point *Y*, all the gas has liquefied and increase in pressure does not change the volume of the gas and the isothermal runs parallel to the vertical axis. (i.e. along *YZ*)

Region *ABC* is referred to as the **liquefaction zone**. In this zone, liquid and gaseous carbon dioxide are in equilibrium.

Question: The diagram below shows isothermals of a gas.



- What is the critical temperature of a gas?
- Which isothermal almost represents the behaviour of an ideal?
- State the condition necessary for liquefaction of a gas.
- Name points *A* and *B*

Relative Atomic Mass and Molecular mass.

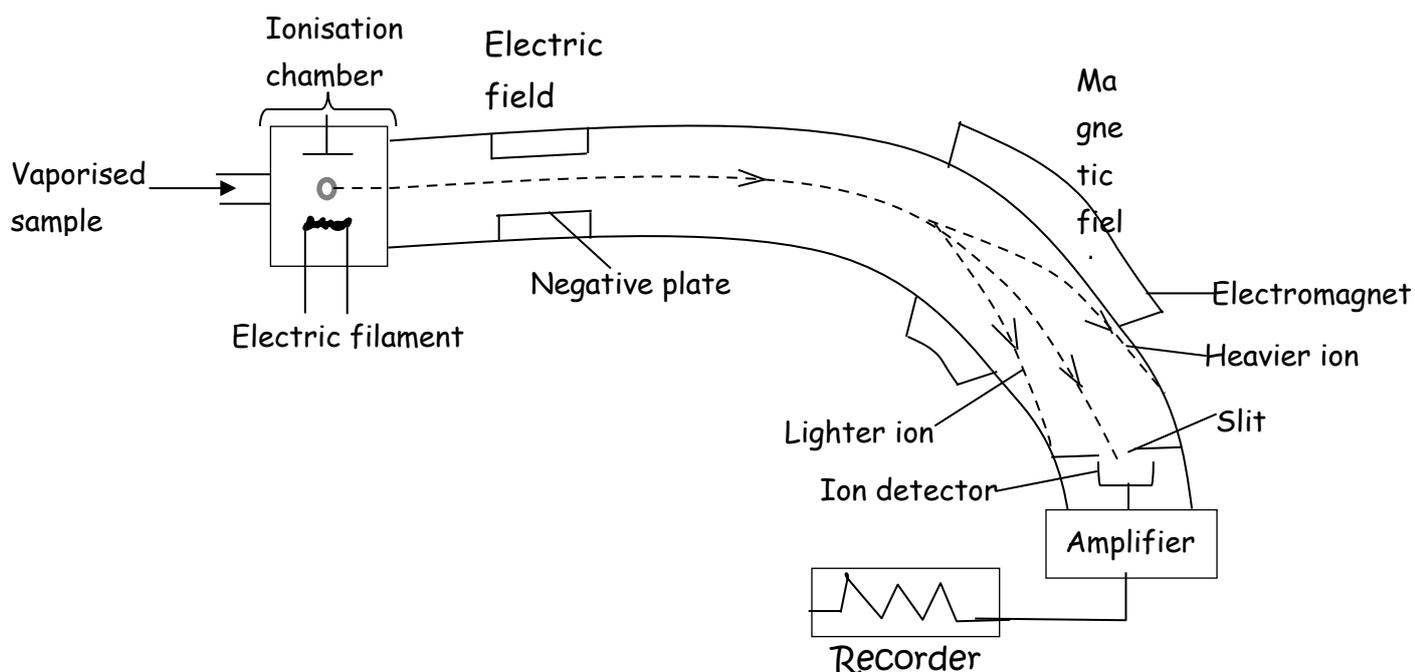
Relative Atomic Mass (R.A.M) is the ratio of the mass of one atom of an element to $\frac{1}{12}$ of the mass of one atom of Carbon-12 isotope.

Relative Molecular Mass (R.M.M) is the ratio of the mass of one molecule of a substance to $\frac{1}{12}$ of the mass of one atom of Carbon-12 isotope.

Both **R.A.M** and **R.M.M.** have no units since they are ratios of similar quantities.

Determination of Relative Atomic Masses/ Molecular Masses Using a mass spectrometer.

Structure



The mass spectrum is evacuated and the sample is vapourised by heating.

The vapourised sample is allowed into the ionization chamber where it is bombarded by fast moving electrons emitted by the electric filament. The gas atoms then become positive ions.

The positive ions are accelerated by the electric field of varying potential so that they enter the magnetic field with the same velocity (or kinetic energy).

The magnetic field deflects the ions into a circular path according to their mass to charge ratio. Ions with high mass to charge ratio deflected less than those with low mass to charge ratio. The electric and magnetic fields are slowly adjusted so that ions of varying mass to charge ratio are received on the detector.

On striking the detector, the ions produce electric current whose magnitude is proportional to the abundance of the isotope from which the ions were formed.

The recorder records currents as a mass spectrum consisting of peaks and relative atomic masses of the different isotopes of the element that formed the ions. The height of the peak corresponds to the relative abundance/intensity of the isotope.

The Relative Atomic Mass (*RAM*) of the element is then calculated as follows;

$$RAM = \sum \left(\frac{\text{Percentage relative abundance}}{100} \times \text{Isotopic mass} \right)$$

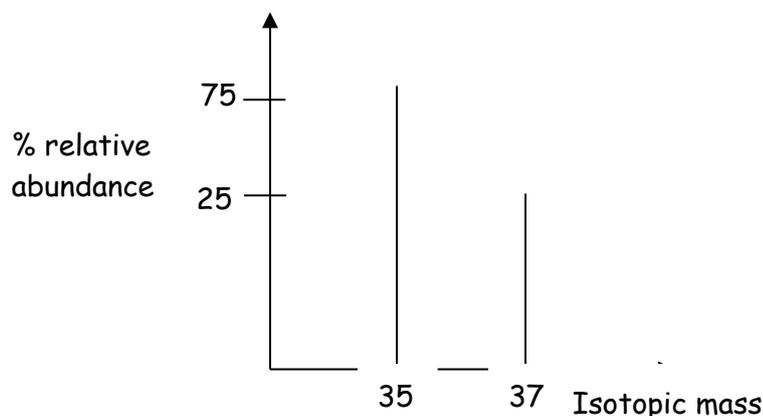
Note:

Relative abundance is the proportion in which an isotope of an element exists naturally as compared to other isotopes of the same element.

Isotopes are atoms of the same element with same atomic number but different number of neutrons hence different atomic masses.

Example:1

Chlorine atom exhibits the following mass spectrum;



Calculate the relative atomic mass of Chlorine.

Solution

$$\begin{aligned} \text{RAM} &= \Sigma \left(\frac{\text{Percentage relative abundance}}{100} \times \text{Isotopic mass} \right) \\ &= \left(\frac{75 \times 35}{100} \right) + \left(\frac{25 \times 37}{100} \right) \\ &= \frac{2625 + 925}{100} \\ &= 35.5 \end{aligned}$$

Therefore, the RAM of chlorine is 35.5

Example: 2

A mass spectrum of chlorine gas shows peaks corresponding to 70, 72, and 74 mass units having relative intensities 9:6:1.

- What particles are responsible for the three peaks?
- What are the relative intensities 9:6:1?
- Explain why the mass spectrum of chlorine gives three molecular peaks yet it is a diatomic gas.

Solution

a) Chlorine has two isotopes of masses 35 and 37 and ^{35}Cl is the most abundant. The peak corresponding to mass unit 70 is due to $^{35}\text{Cl}-^{35}\text{Cl}$, the peak at 72 is due to $^{35}\text{Cl}-^{37}\text{Cl}$ and the last peak at 74 is due to $^{37}\text{Cl}-^{37}\text{Cl}$. Chlorine molecules formed from two identical atoms of

b) The relative intensities in terms of percentage abundances

$$\text{Total ratio} = 9 + 6 + 1 = 16$$

$$\% \text{ abundance of peak at } 70 = \frac{9}{16} \times 100 = 56.25\%$$

$$\% \text{ abundance of peak at } 72 = \frac{6}{16} \times 100 = 37.5\%$$

$$\% \text{ abundance of peak at } 74 = \frac{1}{16} \times 100 = 6.25\%$$

c) Chlorine has two isotopes ^{35}Cl and ^{37}Cl and the three peaks in mass spectrum are due to chlorine molecules when two atoms of ^{35}Cl combine, one atom of ^{35}Cl and one atom of ^{37}Cl combine and the last peak is when two atoms of ^{37}Cl combine.

Example: 3

Copper has a relative mass of 63.55 and consists of two isotopes of mass number 63 and 65. Calculate the percentage of the isotopes.

Let the % abundance of isotope of mass number 63 be m

\Rightarrow the % abundance of isotope of mass number 65 = $(100 - m)$

$$\text{Recal: RAM} = \Sigma \left(\frac{\text{Percentage relative abundance}}{100} \times \text{Isotopic mass} \right)$$

$$63.55 = \left(\frac{63 \times m}{100} \right) + \left(\frac{65(100 - m)}{100} \right)$$

$$6355 = 63m + 65(100 - m)$$

$$65m - 63m = 6500 - 6355$$

$$m = \frac{145}{2} = 72.5\%, \quad \text{therefore, } (100 - m) = 100 - 72.5 = 27.5\%$$

\therefore the % abundance of the isotopes of masses 63 and 65 are 72.5% and 27.5% respectively

Advantages of using a mass spectrometer over other method in determining relative atomic mass.

- The mass spectrometer gives very accurate results
- It requires a very amount of the sample
- It is fast

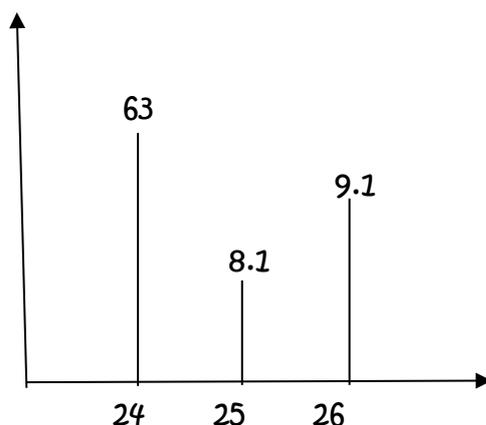
Question:

1. A vapourised sample of element **Q** was introduced into a mass spectrometer and was found to give two isotopes whose masses are 54.999 and 86.9092. It was found that the masses were in a ratio of 5.4: 2.1 respectively. Calculate the RAM of **Q**.
2. Chlorine gas is diatomic that gives a mass spectrum with peaks of masses 70, 72 and 74. The heights of the peaks are in a ratio 9:6:1 respectively.
 - a) Calculate the RMM of chlorine.
 - b) Deduce the RAM of chlorine.
 - c) Calculate the relative abundance of chlorine-35 and chlorine-37.
3. The following information was obtained from the spectrum of lead.

Isotope	Detector current (mA)
204	0.16
206	2.72
207	2.50
208	5.92

Calculate; i) the relative abundance of each isotope
ii) the relative atomic mass of lead

4. Neon has three isotopes, ${}_{10}^{20}\text{Ne}$ (percentage abundance 90.5%); ${}_{10}^{21}\text{Ne}$ (percentage abundance of 0.3%) and ${}_{10}^{22}\text{Ne}$ (percentage abundance of 9.2%). Calculate the mass of Neon.
5. The figure below shows the mass spectrum of magnesium. The heights of the three peaks and mass numbers of the isotopes are shown in the figure below.



Calculate the relative atomic mass of magnesium.

UNEB 2007 P/2 Qn: 1 (a), (b), (c) and (d)

UNEB 2005 P/2 Qn: 1 (b)

UNEB 2004 P/1 Qn: 3 (b)

UNEB 2002 P/1 Qn: 2

Determination of relative molecular masses of gases and volatile liquids.

a) **Determination of relative molecular masses of gases**

Procedure:

- A container filled with pure and dry gas is weighed to determine the mass of the container and gas.
- The container is evacuated and weighed again.
- The volume of the container is found by filling the container with water and measuring the volume of water using the measuring cylinder.

- Weighing is done at **constant temperature (T) and pressure (P)** that are also recorded.

Treatment of results

Let the mass of container + dry gas = ag

Let the mass of empty container = bg

Therefore, mass of gas = $(a-b) = mg$

The volume of the container = Volume of gas = Vg

From Ideal gas equation, $PV = nRT$

Therefore, $M_r = \frac{mRT}{PV}$ where M_r = relative molecular mass of the gas

Alternatively: The volume of a gas is calculated at s.t.p. and then the mass of one mole of the gas determined. This mass is RMM of the gas. i.e. find the mass of $22.4l$ of the gas at s.t.p. which gives its RMM

b) Determination of relative molecular masses of volatile liquids

Procedure

A known mass (mg) of a volatile liquid is injected into a **gas syringe**. The vapour evaporates and moves the piston. When the piston becomes steady stable, the volume (V) of the vapour is recorded. The temperature (T) around the syringe and atmospheric pressure (P) are recorded.

Treatment of results

From ideal gas equation, $PV = nRT$

Therefore, $M_r = \frac{mRT}{PV}$ where M_r = relative molecular mass of the gas

Questions:

1. When $0.16g$ of a volatile liquid X was injected into a gas syringe at $120^\circ C$ and $745mmHg$ pressure, $51.5cm^3$ of the liquid vapour was produced. Calculate the formula mass of X . The molar volume of a gas is $22.4 dm^3$ at s.t.p.
2. A gas syringe containing $20.6cm^3$ of air at $48^\circ C$ was injected with $0.246g$ of a volatile liquid. The volume of the gas in the syringe at $48^\circ C$ was found to be $63.7cm^3$ at a pressure 1.01×10^5 pa. Calculate the molar mass of the liquid.

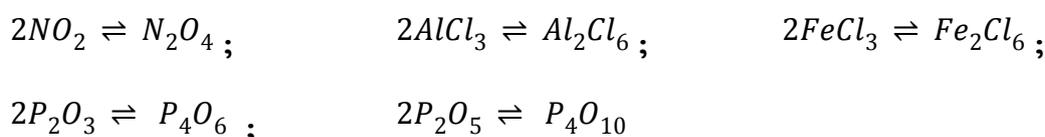
Abnormal values of relative molecular masses

Sometimes RMM determined from the experiments above are not in agreement with the real/actual values expected from the formula. This may be due to;

a) Association of a compound

Association is the combination of two or more like molecules to form a single molecule of higher molecular mass.

Association normally occurs when the experiment is carried out at low temperatures and it results in decrease in the number of moles. Examples include;



For example, if n moles of X associate and α is the degree of association, then

	nX	\rightleftharpoons	X_n
<i>Initial moles</i>	n		0
<i>Reaction moles</i>	$n\alpha$		α
<i>Equilibrium moles</i>	$n(1 - \alpha)$		α

Moles before association = Initial moles of X = n moles

*Moles after association = Total equilibrium moles = $n(1 - \alpha) + \alpha$
= $n + \alpha(1 - n)$ moles*

$$\frac{\text{RMM after association}}{\text{RMM before association}} = \frac{\text{Moles before association}}{\text{Moles after association}}$$

$$\frac{\text{RMM after association}}{\text{RMM before association}} = \frac{n}{n + \alpha(1 - n)}$$

$n + \alpha(1 - n)$ RMM after association = n RMM before association

$\alpha(1 - n)$ RMM after asstn = n RMM before association - n RMM after association

$$\alpha = \frac{n \text{ RMM before association} - n \text{ RMM after association}}{(1 - n) \text{ RMM after association}}$$

$$\alpha = \frac{n(\text{RMM before association} - \text{RMM after association})}{(1 - n) \text{ RMM after association}}$$

$$\alpha = \frac{n(\text{RMM after association} - \text{RMM before association})}{(n - 1) \text{ RMM after association}}$$

where n = number of molecules that associated

Example:

If molar mass of aluminium chloride at 192°C is 187, calculate the degree of dimerization of aluminium chloride.

Solution:

Method I

Molar mass of AlCl_3 after dimerization = 187 g

Molar mass of AlCl_3 before dimerization = $27 + (3 \times 35.5) = 133.5$ g

n = number of molecules for $\text{AlCl}_3 = 2$

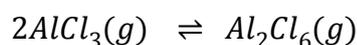
$$\text{Using } \alpha = \frac{n(\text{RMM after association} - \text{RMM before association})}{(n - 1) \text{ RMM after association}}$$

$$\alpha = \frac{2(187 - 133.5)}{(2 - 1) \times 187} = 0.572 = 57.2\%$$

Method II

Molar mass of AlCl_3 after dimerization = 187 g

Molar mass of AlCl_3 before dimerization = $27 + (3 \times 35.5) = 133.5$ g



Initial moles 2 0

Reaction moles 2α α

Equilibrium moles $2(1 - \alpha)$ α

Moles of AlCl_3 before dimerization = 2 moles

Moles of AlCl_3 after dimerization = $2(1 - \alpha) + \alpha = 2 - \alpha$ moles

$$\frac{\text{RMM after dimerization}}{\text{RMM before dimerization}} = \frac{2}{2 - \alpha}$$

$$\frac{187}{133.5} = \frac{2}{2 - \alpha} \quad ; \quad \alpha = \frac{(2 \times 187) - (2 \times 133.5)}{187} = 0.572 = 57.2\%$$

b) Dissociation of a compound

Dissociation is the breaking up of a molecule of a compound into other smaller molecules.

Examples of compounds that dissociate include:



The RMM of such compounds determined experimentally will be smaller than the actual/expected value.

Degree of dissociation (α) is the fraction or percentage per mole of the original substance that has dissociated into simpler substances.

For example, if n molecule of AB dissociates into A and B and α is the degree of dissociation of AB .



Initial moles n 0 0

Reaction moles $n\alpha$ $n\alpha$ $n\alpha$

Equilibrium moles $n(1 - \alpha)$ $n\alpha$ $n\alpha$

Moles before dissociation = Initial moles = n moles

Moles after dissociation = Total equilibrium moles = $n(1 - \alpha) + n\alpha + n\alpha$

$$= n(1 + \alpha) \text{ moles}$$

$$\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} = \frac{\text{Moles after dissociation}}{\text{Moles before dissociation}}$$

$$\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} = \frac{n(1 + \alpha)}{n}$$

$$n(1 + \alpha) \text{ RMM after dissociation} = n \text{ RMM before dissociation}$$

$$(1 + \alpha) = \frac{n \text{ RMM before dissociation}}{n \text{ RMM after dissociation}}$$

$$\alpha = \left(\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} \right) - 1$$

Example:

Dinitrogen tetraoxide dissociates partly to nitrogen (IV) oxide according to the equation;

$$N_2O_4 \rightleftharpoons 2NO_2$$

The RMM of dinitrogen tetraoxide at 27°C was found to be 76.6. Find the degree of dissociation.

Method I

$$\text{RMM of } N_2O_4 \text{ before dissociation} = (2 \times 14) + (4 \times 16) = 92 \text{ g}$$

$$\text{RMM of } N_2O_4 \text{ after dissociation} = 76.6 \text{ g}$$

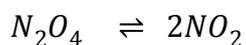
$$\alpha = \left(\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} \right) - 1$$

$$\alpha = \frac{92}{76.6} - 1 = 0.201 = 20.1\%$$

Method II

$$\text{RMM of } N_2O_4 \text{ before dissociation} = (2 \times 14) + (4 \times 16) = 92 \text{ g}$$

$$\text{RMM of } N_2O_4 \text{ after dissociation} = 76.6 \text{ g}$$



Initial moles	1	0
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Reaction moles	α	2α
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Equilibrium moles	$(1 - \alpha)$	2α
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Moles before dissociation = 1 moles

Moles after dissociation = $(1 - \alpha) + 2\alpha = (1 + \alpha)$

$$\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} = \frac{\text{Moles after dissociation}}{\text{Moles before dissociation}}$$

$$\frac{\text{RMM before dissociation}}{\text{RMM after dissociation}} = \frac{(1 + \alpha)}{1} \quad ; \quad \frac{92}{76.6} = \frac{(1 + \alpha)}{1}$$

$$\alpha = \frac{92}{76.6} - 1 = 0.201 = 20.1\%$$

Trial Questions:

1. State Graham's law of diffusion.
 - a) (i) State Graham's law of diffusion.
 - (ii) The ratio of the rate of diffusion of a gas Y to that of nitrogen was found to be 0.366. Calculate the relative molecular mass of Y.
2. 6.3cm³ of a gas X diffuses through a porous partition in 5 minutes. If 7.4cm³ of oxygen diffuses through the same porous partition in the same time, calculate the molecular mass of X.
3. A hydro carbon P, contains 85.7% of carbon. If oxygen diffuses 1.31 times faster than P, calculate the;
 - a) Empirical formula of P
 - b) Relative molecular mass of P
 - c) Molecular formula of P
4. Nitrogen diffuses 1.29 times faster than a gas Y. Calculate the RMM of Y.
5. Nickel forms a carbonyl Ni(CO)_n, if carbon monoxide diffuses 2.46 times faster than the carbonyl. Calculate the
 - i) molecular mass of the carbonyl
 - ii) value of n (given that Ni=59, C=12, O=16)
6. The rate diffusion of gas A to B is in the ratio of 1:2. If the RMM of A is 16. Calculate the RMM of B.
7. If the volume of a gas collected at 60°C and 1.05×10⁵ Nm⁻² is 60cm³. What would be the volume of the gas at s.t.p.?

8. State Dalton's law of partial pressures and show it may be derived from the ideal gas equation.
9. A mixture of oxygen and nitrogen in which the mole fraction of oxygen is 0.25, is at a pressure of 1.0 atm. Calculate the partial pressure of each gas in the mixture (assumed to be ideal)
10. 12g of nitrogen, 0.4g of hydrogen and 9g of oxygen were put in a 1 litre vessel at a pressure of 22.4 atm. Calculate the partial pressures of the individual gases.
11. A mass spectrum of dichloromethane shows peaks corresponding to 84, 86 and 88 mass units having relative intensities of 9:6:1 respectively.
 - a) What particles are responsible for the three peaks?
 - b) What are the relative intensities of 9:6:1
 - c) Explain why the mass of chlorine gives three peaks while that of fluorine gives only one molecular peak.
12. Phosphorous (V) chloride dissociates when heated to a temperature of 165°C to give Phosphorous (III) chloride and chlorine gas. If its molar mass 172, calculate the degree of dissociation.