UNIT 4 - STATES OF MATTER

There are three different states of matter: solid, liquid, and gaseous. The arrangement of particles is other in each state:

	SOLID	LIQUID	GASEOUS
The relative position of particles	Particles are tightly packed in fixed arrangement crystalline structure	Irregular packing of particles (clusters). Particles slide over each other.	Particles are far apart and randomly spaced. There is no fixed arrangement.
Relative movement of particles	Particles vibrate about a fixed position	Particles slide over each other – translatory movement	Rapid and random movement – frequent collisions
Intermolecular Forces	Very strong IMF	Relatively weak IMF	Very weak IMF

The Kinetic Theory of Gases

Origin of pressure – pressure can be explained by the Kinetic Theory as <u>arising from the force exerted by molecules/atoms impacting on the walls of a container</u>.

The kinetic theory of gases is basically the idea that molecules in gases are constantly moving. There are 5 main assumptions of this theory:

1) Gas molecules move rapidly and randomly
2) The distance between the gas molecules is much greater than the diameter of the molecules – the volume of the molecules is <u>negligible</u>
3) There are no forces of attraction or repulsion between the molecules
4) All collisions between the molecules are elastic, meaning no kinetic energy is lost in collisions
5) The temperature of the gas is related to the average kinetic energy of the molecules: Kinetic Energy ∝ Temperature
    - The faster the molecules move, the more kinetic energy they have, the higher temperature the substance has
    - Absolute temperature is based on the idea that at some temperature, all kinetic motion will stop. This is known as absolute zero.

Ideal Gases

A theoretical gas that fits the description of the assumptions is an ideal gas. The ideal gas equation is:

$$PV = nRT$$

P - pressure in Nm^{-1}
V - volume of the gas in m^3
n - number of moles of the gas
R - a constant referred to as the Universal Gas Constant
T - absolute temperature in Kelvins (K)

How to derive the Ideal Gas Equation

Boyle's law (relating volume and pressure), Charles's law (relating volume and temperature), and Avogadro's law (relating volume to the number of moles) may be combined into a single expression relating all four terms.

Avagadro's Law

$V \propto n$
$V = nk$
$k = \frac{V}{n}$

Equal volumes of any ideal gas contain the same number of moles measured under the same conditions of temperature and pressure.

$P \propto \frac{1}{V}$
$P = \frac{k}{V}$
$k = PV$

$\boxed{P_1V_1 = P_2V_2}$

Boyles' Law

The volume of a gas varies inversely with the pressure exerted by the gas if the number of moles and temperature of the gas is held constant; the product of pressure and volume is constant.

$V \propto T$
$V = kT$
$k = \frac{V}{T}$

Charles' Law

The volume of a gas varies directly with the absolute temperature (K) if pressure and number of moles of gas are constant.

1 $k = \frac{V}{n}$, $k = PV$, 3 $k = \frac{V}{T}$

How to find the Mr of a gas using the Ideal Gas Equation

A flask of volume 2.0 dm3 was found to contain 5.28g of gas. The pressure in the flask was 200kPa and the temperature was 20°C. Calculate the relative molecular mass of the gas.
(R = 8.31 JK-1mol-1)

$PV = nRT$

$PV = \frac{m}{Mr} RT$

$Mr = \frac{mRT}{PV}$

1. vol = $2.0 \times 10^{-3} m^3$.
2. mass = 5.28g. $(n = \frac{mass}{Mr})$
3. P = 200kPa $\longrightarrow 2.0 \times 10^5 Nm^{-2}$.
4. T = 20 + 273 = 293K.

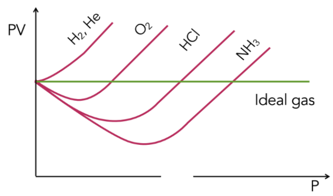
$$\frac{5.28 \times 8.31 \times 293}{(2.0 \times 10^5) \times (2.0 \times 10^{-3})} = 32.1 \ g mol^{-1}$$

Limitations of the Ideal Gas Laws

Scientists have taken accurate measurements to show how the volumes of gases change with temperature and pressure. These show us that gases do not always behave exactly as we expect an ideal gas to. This is because real gases do not always obey the kinetic theory in two ways:

- There is not "zero attraction" between the molecules
- We cannot ignore the volume of the molecules themselves.

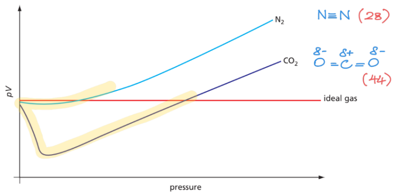
Real Gas vs. Ideal Gas Laws



The validity of Boyles' Law was tested over a wide range of pressures and it was found that none of the gases obeyed the law.

If the gases obeyed Boyles law, the plots of PV against P should be parallel to the x-axis.

High Pressure:
The main reason why gases behave less ideally at high pressures is because the volume of their molecules become increasingly significant proportion of the overall volume of the gas.



Low Temperatures:

The main reason why gases behave less ideally at low temperatures is because the intermolecular forces of attraction become comparable in size to the bouncing forces the molecules experience. This causes the collisions between the molecules to be inelastic.

Deviation in Ideal Gas Behaviour

The deviation from ideal behaviour is expressed by introducing a factor Z known as compressibility factor in the ideal gas equation. Z may be expressed as Z = PV / nRT

- In case of ideal gas, PV = nRT ∴ Z = 1
- In case of real gas, PV ≠ nRt ∴ Z ≠ 1

Thus in case of real gases Z can be < 1 or > 1

- When Z < 1, it is a negative deviation. It shows that the gas is more compressible than expected from ideal behaviour.
- When Z > 1, it is a positive deviation. It shows that the gas is less compressible than expected from ideal behaviour.

The causes of deviations from ideal behaviour may be due to the following two assumptions of kinetic theory of gases.

- The volume occupied by gas molecules is negligibly small as compared to the volume occupied by the gas.
- The forces of attraction between gas molecules are negligible.

The first assumption is valid only at low pressures and high temperature, when the volume occupied by the gas molecules is negligible as compared to the total volume of the gas. But at low temperature or at high pressure, the molecules being in compressible the volumes of molecules are no more negligible as compared to the total volume of the gas.