

4 States of Matter

The study of the particles in solids, liquids and gases and the interactions between them is important in understanding the physical properties of substances.

4.1 The gaseous state: ideal and real gases and $pV = nRT$



4.2 The liquid state



STATES OF MATTER

www.megalecture.com

MEGA LECTURE

For Live Online Classes
megalecture@gmail.com

STATES OF MATTER

www.youtube.com/megalecture

4 States of matter

The study of the particles in solids, liquids and gases and the interactions between them is important in understanding the physical properties of substances.

Learning outcomes

Candidates should be able to:

4.1 The gaseous state: ideal and real gases and $pV = nRT$

- a) state the basic assumptions of the kinetic theory as applied to an ideal gas
- b) explain qualitatively in terms of intermolecular forces and molecular size:
 - (i) the conditions necessary for a gas to approach ideal behaviour
 - (ii) the limitations of ideality at very high pressures and very low temperatures
- c) state and use the general gas equation $pV = nRT$ in calculations, including the determination of M_r

4.2 The liquid state

- a) describe, using a kinetic-molecular model, the liquid state, melting, vaporisation, vapour pressure

KINETIC THEORY OF GASES

The idea that molecules in gases are in constant movement is called the kinetic theory of gases. This theory makes certain assumptions:

1. Gases are made up of tiny particles, “molecules” in a state of rapid, random motion.
2. Average kinetic energy of particles is directly proportional to the temperature of gas in Kelvin
3. All collisions are perfectly elastic i.e. there is no loss of kinetic energy.
4. Collisions between the molecules and the walls of the container give rise to pressure.

1

KINETIC THEORY OF GASES

5. The volume of molecules is negligible compared to the volume of the gas.
(The distance between the gas molecules is much greater than the diameter of the molecules so the volume of the molecules is negligible)
6. Attractive forces between particles are negligible.

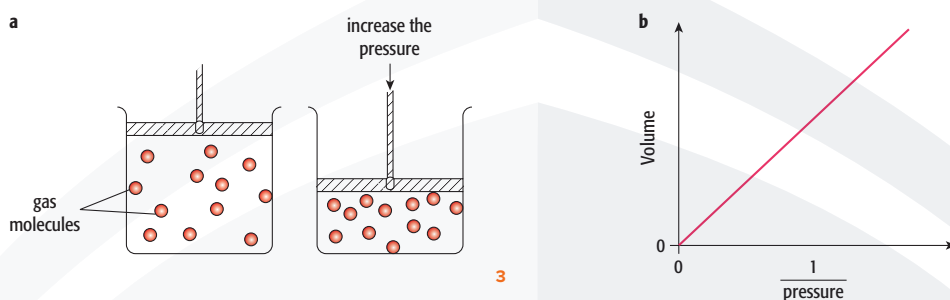
2

BOYLE'S LAW

The volume of a gas varies inversely with the pressure exerted by the gas if the number of moles and temperature of gas are held constant. This relationship is known as Boyle's law. The product of pressure (P) and volume (V) is a constant:

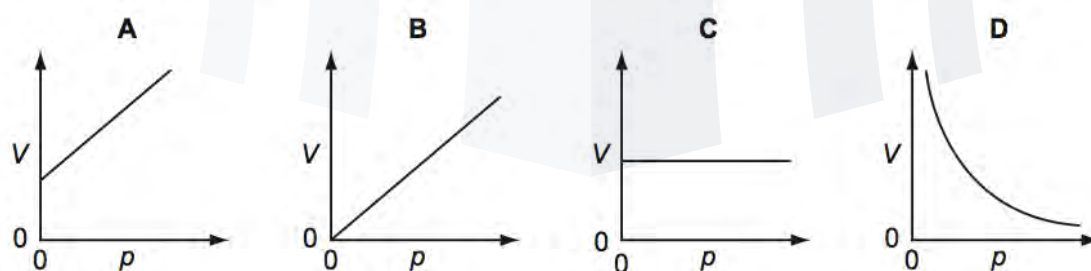
$$PV = k \text{ (constant)}$$

$$P_1V_1 = P_2V_2$$



SKILL CHECK 1

Which diagram shows the correct graph of V against p for a fixed mass of an ideal gas at constant temperature?



4

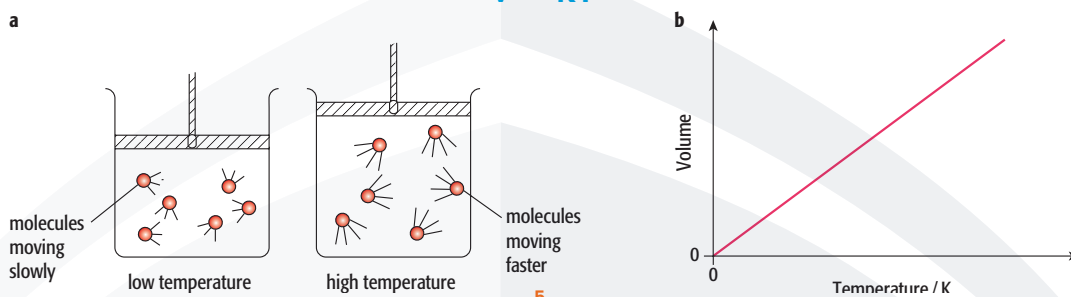
CHARLES' LAW

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

Mathematically, the ratio of volume (V) and temperature (T) is a constant:

$$V/T = k \text{ (constant)}$$

$$V = kT$$



5

AVAGADRO'S LAW

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

Mathematically, the ratio of volume (V) to number of moles (n) is a constant:

$$V/n = k \text{ (constant)}$$

$$V = nk$$

6

IDEAL GAS LAW

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.

This expression is the ideal gas law:

$$PV = nRT$$

P = pressure of the gas in Nm^{-2} n = number of moles of the gas

V = volume of the gas in m^3 T = absolute temperature of the gas in K

R = a constant referred to as the Universal gas constant*

*Unit and value of $R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$

7

SKILL CHECK 2

The general gas equation can be used to calculate the M_r value of a gas.

For a sample of a gas of mass m g, which expression will give the value of M_r ?

A $M_r = \frac{mpV}{RT}$ B $M_r = \frac{pVRT}{m}$ C $M_r = \frac{mRT}{pV}$ D $M_r = \frac{pV}{mRT}$

8

SKILL CHECK 3

The volume of a sample of ammonia is measured at a temperature of 60°C and a pressure of 103 kPa . The volume measured is $5.37 \times 10^{-3}\text{ m}^3$.

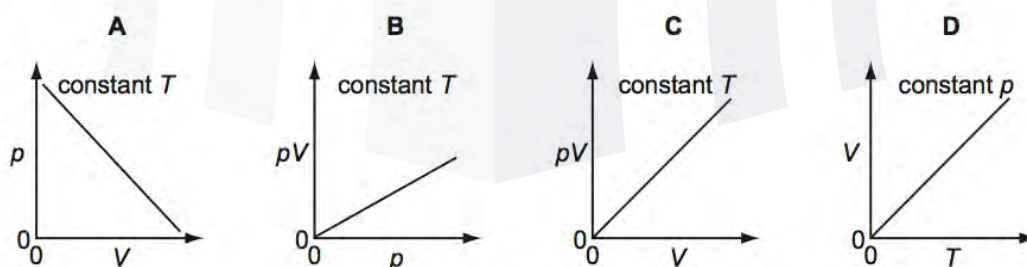
What is the mass of the sample of ammonia, given to two significant figures?

- A** 0.00019 g **B** 0.0034 g **C** 0.19 g **D** 3.4 g

9

SKILL CHECK 4

Which diagram correctly describes the behaviour of a fixed mass of an ideal gas? (T is measured in K.)

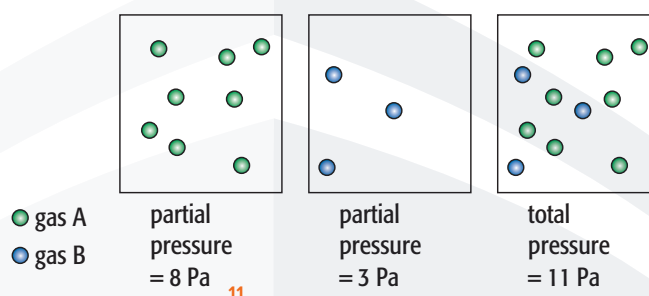


10

DALTON'S LAW OF PARTIAL PRESSURES

A mixture of gases exerts a pressure that is the sum of the pressures that each gas would exert if it were present alone under the same conditions.

Consider two gases A and B occupying a given volume individually at a given temperature. Let their pressures be $p_a = 8 \text{ Pa}$ and $p_b = 3 \text{ Pa}$. Now consider the same amount of the two gases A and B occupying the same volume at the same temperature all together.



DALTON'S LAW OF PARTIAL PRESSURES

Let the total pressure be P_T . Then according to Dalton's law:

$$P_T = p_a + p_b$$

p_a and p_b are referred to as partial pressures

Hence $p = x P$ where p is the partial pressure, and x is the mol fraction of the gas

SKILL CHECK 5

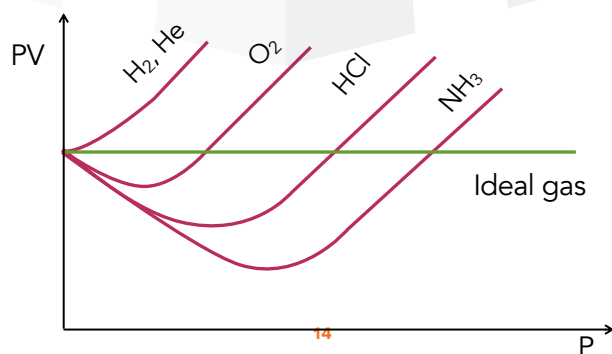
- a.** 80 g of oxygen, 140 g of nitrogen, 44 g of carbon dioxide and 60 g of argon when confined in a vessel exerted pressure of a $3.6 \times 10^6 \text{ Nm}^{-2}$. Hence calculate the partial pressure of each of the gases.
- b.** A mixture of gases at a pressure $7.50 \times 10^4 \text{ Nm}^{-2}$ has the volume composition 40% nitrogen, 35% oxygen and 25% carbon dioxide. What is the partial pressure of each gas?
- c.** 2.0 dm^3 of nitrogen at a pressure of $1.0 \times 10^5 \text{ Pa}$ and 5.0 dm^3 of hydrogen at a pressure of $5.0 \times 10^5 \text{ Pa}$ are injected into a 10 dm^3 vessel. What is the pressure of the mixture of the gases?

13

REAL GASES VS IDEAL GAS

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

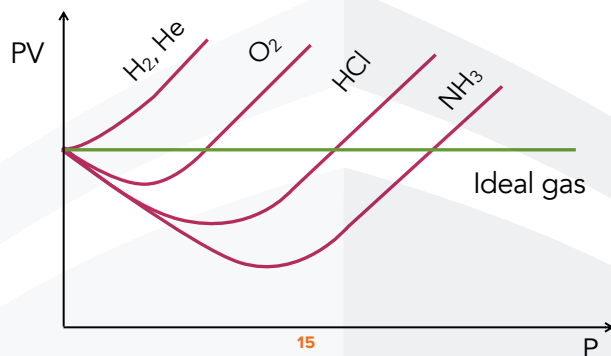
If the gases obeyed Boyle's law, the plots of PV against P should be parallel to the x axis. None of the gases gave such a graph. Two types of deviations are seen.



REAL GASES VS IDEAL GAS

All gases except hydrogen and helium, first showed a minimum, and then increased with increasing pressure.

Hydrogen and helium never showed a minimum, always increasing with increasing pressure. (+ve deviation).



REAL GASES VS IDEAL GAS

The hypothetical gas, which would strictly obey the gas laws, is called the ideal gas or perfect gas, in contrast to all known gases which are referred to as real gases. Ideal gas is one that obeys the gas laws/ $pV = nRT$ under all conditions of temperature and pressure

These show us that gases do not always behave exactly as we expect an ideal gas to behave. This is because real gases do not always obey the kinetic theory in two ways:

1. There is not zero attraction between the molecules
2. We cannot ignore the volume of the molecules themselves.

16

REAL GASES VS IDEAL GAS

These differences are especially noticeable at very high pressures and very low temperatures. Under these conditions:

- The molecules are close to each other
- The volume of the molecules is not negligible compared with the volume of the container
- There are Van der Waals' forces of attraction between the molecules
- Attractive forces pull the molecules towards each other and away from the walls of the container
- The effective volume of the gas is smaller than expected for an ideal gas.

17

REAL GASES VS IDEAL GAS

In hydrogen and helium the intermolecular attractions are negligible and hence they do not get compressed more than expected, and hence they do not show a minimum.

The two assumptions would become true only when the volume occupied by the gas tends to infinity. Volume would tend to infinity when pressure decreases and temperature increases.

Thus real gases tend to be ideal i.e tend to obey the gas laws at low pressures and high temperatures.

Conversely at very high pressures and very low temperatures the laws become invalid.

18

SKILL CHECK 6

Which would behave the **least** like an ideal gas at room temperature?

- A carbon dioxide
- B helium
- C hydrogen
- D nitrogen

19

THE LIQUID STATE

The particles in a liquid are still close to each other but have enough kinetic energy to keep sliding past each other in fairly random way, but not as freely as gas particles do.

For brief periods, the particles in liquids are arranged in slightly ordered way. but this order is always broken up when the particles gain kinetic energy from neighbouring particles.

20

THE LIQUID STATE

When we heat a liquid:

Energy transferred to the liquid makes the particles move faster.

Forces of attraction between particles weaken.

Particles with most energy are the first to escape from forces holding them together.

The liquid evaporates - this happens at temperatures below the boiling point.

21

THE LIQUID STATE

Forces weaken enough for all particles to become free from each other.

The liquid boils.

This change is called **vaporisation**.

The energy required to change one mole of liquid to one mole of gas is called enthalpy change of vaporisation.

22

THE LIQUID STATE

When we cool a vapour, the particles:

Lose kinetic energy so the molecules move around less quickly.

Experience increasing forces of attraction.

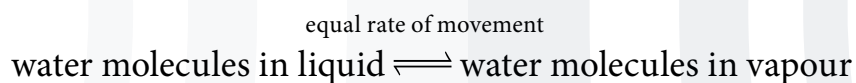
Move more slowly and become closer. The gas liquefies.

This change of state is called **condensation**. These changes in state are reversible. These changes involve opposite energy transfers.

23

THE LIQUID STATE

At equilibrium the concentration of water molecules in the vapour remains constant.



Pressure exerted by a vapour in equilibrium with its liquid is called vapour pressure.

The temperature at which vapour pressure is equal to the atmospheric pressure is the boiling point of the liquid.

24