HL IB Chemistry

Your notes

Ideal Gases

Contents

Your notes

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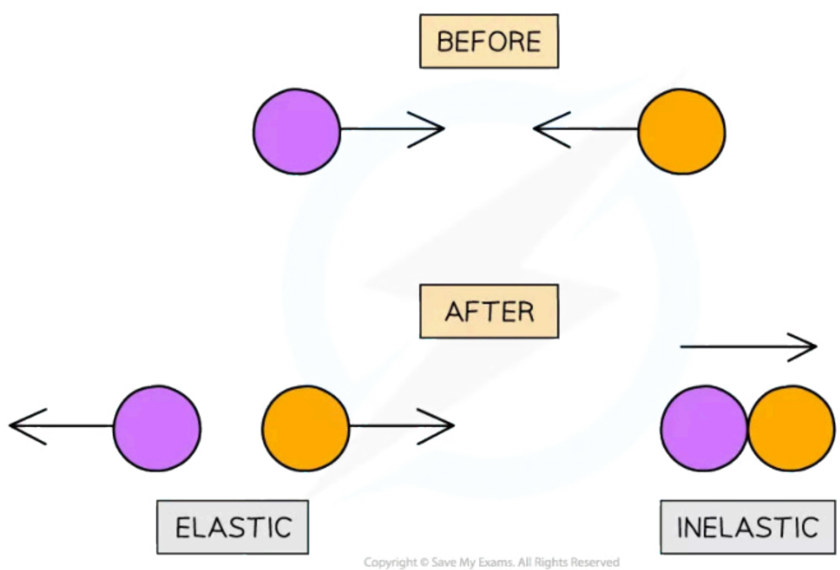
Ideal Gases

- **The kinetic theory of gases** states that molecules in gases are constantly moving
- The theory makes the following assumptions:
  - The gas molecules are moving very fast and randomly
  - The molecules hardly have any volume
  - The gas molecules do not attract or repel each other (**no intermolecular forces**)
  - No kinetic energy is lost when the gas molecules collide with each other (**elastic collisions**)
  - The temperature of the gas is directly proportional to the average kinetic energy of the molecules
- Gases that follow the kinetic theory of gases are called **ideal gases**
- However, in reality gases do not fit this description exactly **but** may come very close and are called **real gases**
- The volume that a gas occupies depends on:
  - Its pressure
  - Its temperature

Diagram to show elastic collisions



In an elastic collision, energy is conserved and the particles colliding strike each other then move away in opposite directions whereas in an inelastic collision kinetic energy is not conserved and the particles usually strike and stick together

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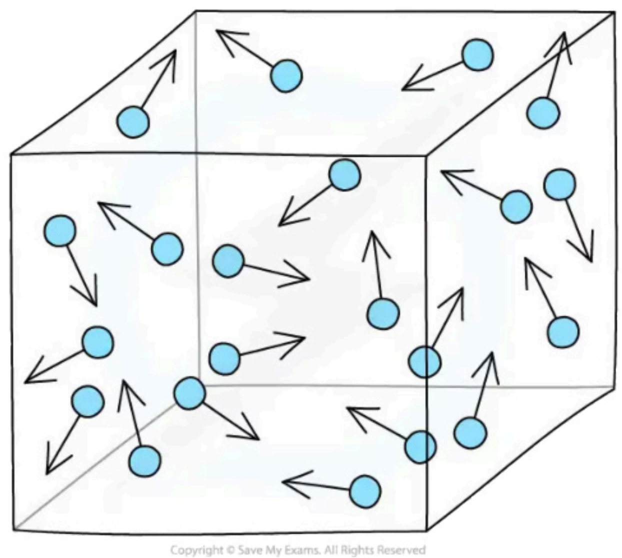
Molar Gas Volume

Molar Gas Volume

- **Gases** in a container exert a **pressure** as the gas molecules are constantly **colliding** with the walls of the container

A particle model of a gas in a container

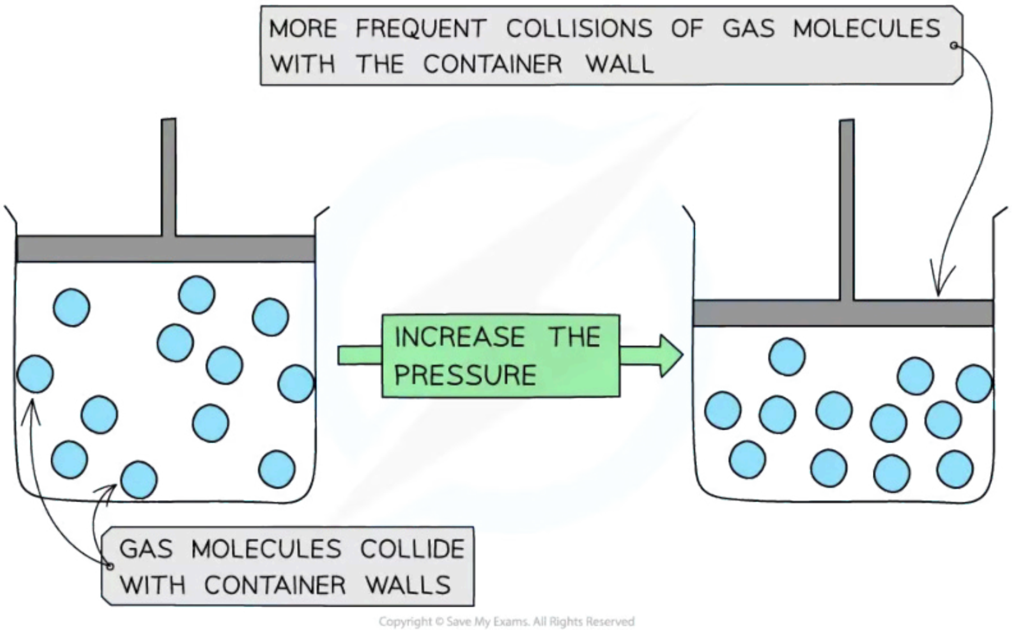


Gas particles exert a pressure by constantly colliding with the walls of the container

Changing gas volume

- **Decreasing** the **volume** (at constant temperature) of the container causes the molecules to be **squashed** together which results in more **frequent** collisions with the container wall
- The **pressure** of the gas **increases**

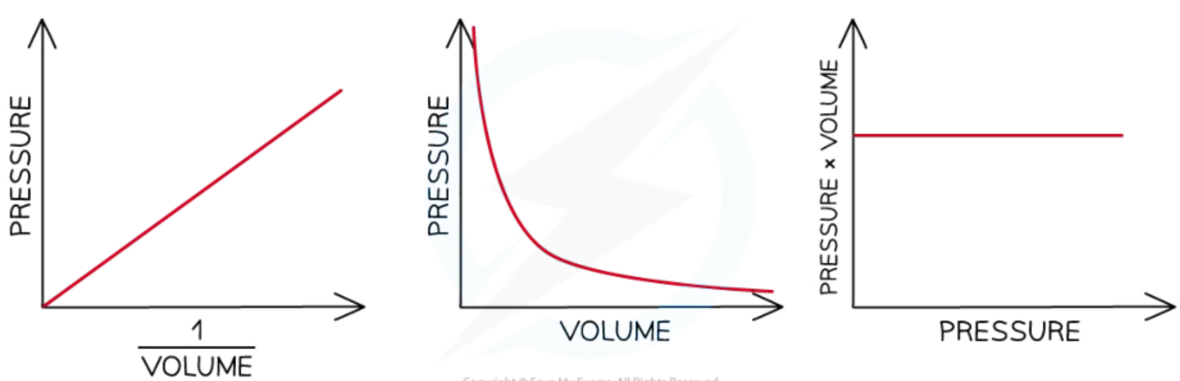
Gas Molecule Collision Frequency with Increasing Pressure Diagram



Decreasing the volume of a gas causes an increased collision frequency of the gas particles with the container wall

- The **pressure** is therefore **inversely proportional** to the **volume** (at constant temperature)
- This is known as **Boyle's Law**
- Mathematically, we say $P \propto 1 / V$ or PV **= a constant**
- We can show a graphical representation of **Boyle's Law** in three different ways:
  - A graph of pressure of gas plotted against 1 / volume gives a straight line
  - A graph of pressure against volume gives a curve
  - A graph of PV versus P gives a straight line

Sketch graphs of Boyles' Law



Three graphs that show Boyle's Law

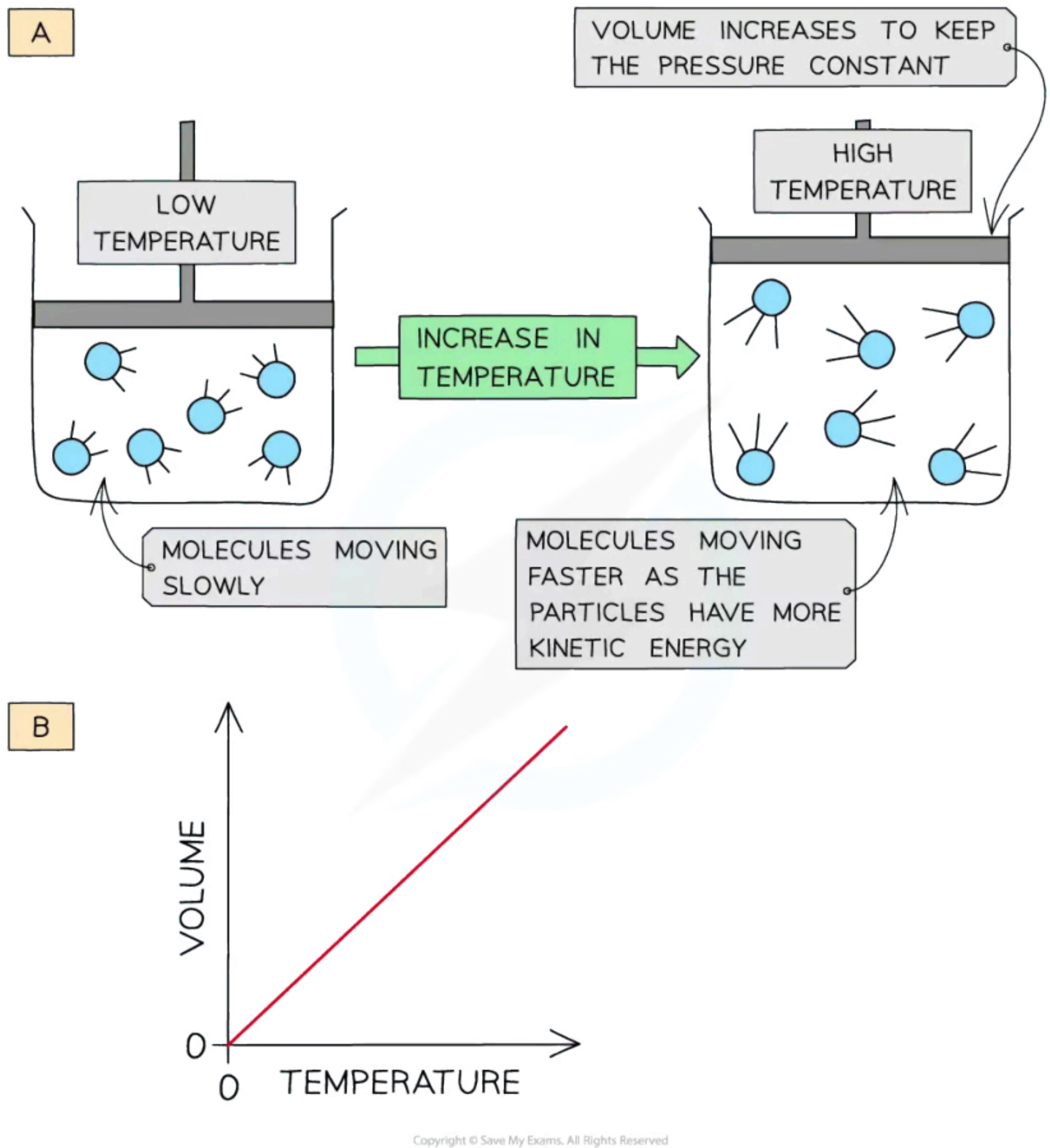
Changing gas temperature

- When a gas is **heated** (at constant pressure) the particles gain more **kinetic energy** and undergo more **frequent collisions** with the container walls
- To keep the **pressure constant**, the molecules must get further apart and therefore the **volume increases**
- The **volume** is therefore **directly proportional** to the **temperature in Kelvin** (at constant pressure)
- This is known as **Charles' Law**
- Mathematically, $V \propto T$ or $\boldsymbol{V / T}$ **= a constant**
- A graph of **volume** against **temperature in Kelvin** gives a straight line

Gas Molecule Collision Frequency with Increasing Temperature Diagram

Your notes

Your notes



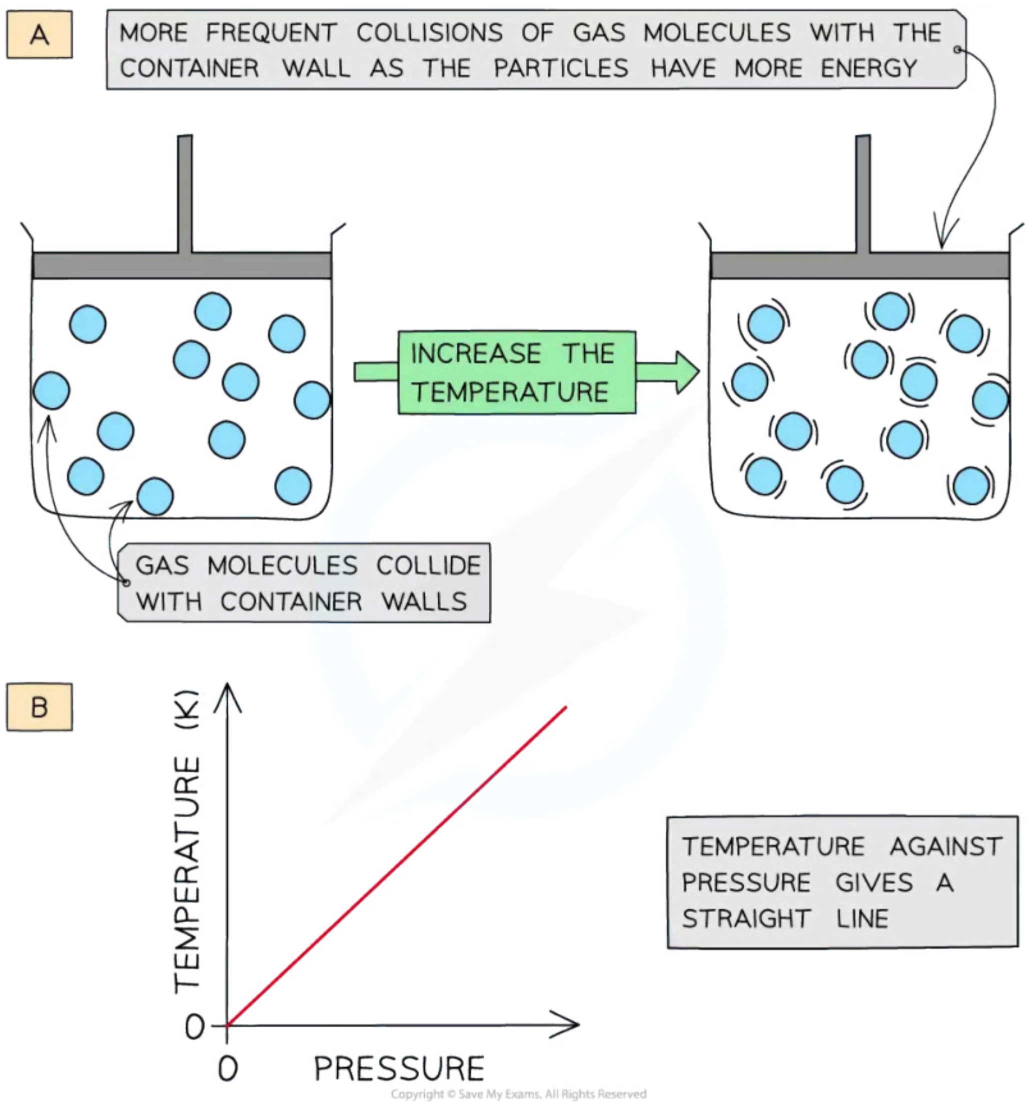
Increasing the temperature of a gas causes an increased collision frequency of the gas particles with the container wall (a); volume is directly proportional to the temperature in Kelvin (b)

Changing gas pressure

- **Increasing** the **temperature** (at constant volume) of the gas causes the molecules to gain more **kinetic energy**
- This means that the particles will move **faster** and **collide** with the container walls more **frequently**
- The **pressure** of the gas **increases**
- The **temperature** is therefore **directly proportional** to the **pressure** (at constant volume)
- Mathematically, we say that $P \propto T$ or $\boldsymbol{P / T}$ **= a constant**

- A graph of **temperature in Kelvin** of a gas plotted against **pressure** gives a straight line

Gas Molecule Collision Frequency with Increasing Pressure at Constant Volume Diagram



Increasing the temperature of a gas causes an increased collision frequency of the gas particles with the container wall (a); temperature is directly proportional to the pressure (b)

Pressure, volume and temperature

- Combining these three relationships together:
  - PV = a constant
  - V / T = a constant
  - P / T = a constant
- We can see how the **ideal gas equation** is constructed

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- PV / T = a constant
- PV = a constant x T
- This constant is made from two components, the number of **moles**, **n**, and the **gas constant**, ***R***, resulting in the overall equation:
  - ***PV = nRT***

Changing the conditions of a fixed amount of gas

- For a fixed amount of gas, ***n*** and ***R*** will be constant, so if you change the conditions of a gas we can ignore ***n*** and ***R*** in the **ideal gas equation**
- This leads to a very useful expression for problem solving

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

- Where P_1, V_1 and T_1 are the initial conditions of the gas and P_2, V_2 and T_2 are the final conditions

Worked example

At 25 °C and 100 kPa a gas occupies a volume of 20 dm^3. Calculate the new temperature, in °C, of the gas if the volume is decreased to 10 dm^3 at constant pressure.

Answer:

Step 1: Rearrange the formula to change the conditions of a fixed amount of gas. Pressure is constant so it is left out of the formula

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

Step 2: Convert the temperature to Kelvin. There is no need to convert the volume to m^3 because the formula is using a **ratio** of the two volumes

$V_1 = 20\ dm^3$

$V_2 = 10\ dm^3$

$T_1 = 25 + 273 = 298$ K

Step 3: Calculate the new temperature

$$T_2 = \frac{10\ dm^3 \times 298\ K}{20\ dm^3} = 149\ K = -124\ °C$$

Worked example

A 2.00 dm^3 container of oxygen at a pressure of 80 kPa was heated from 20 °C to 70 °C The volume expanded to 2.25 dm^3. What was the final pressure of the gas?

Answer:

Step 1: Rearrange the formula to change the conditions of a fixed amount of gas

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1}$$

Step 2: Substitute in the values and calculate the final pressure

P_1 = 80 kPa

V_1 = 2.00 dm^3

V_2 = 2.25 dm^3

T_1 = 20 + 273 = 293 K

T_2 = 70 + 273 = 343 K

$$P_2 = \frac{80 \text{ kPa} \times 2.00 \text{ dm}^3 \text{ x } 343 \text{ K}}{293 \text{ K} \times 2.25 \text{ dm}^3} = 83 \text{ kPa}$$

Your notes

The Ideal Gas Equation

Your notes

Ideal Gas Equation

- The **ideal gas equation** shows the relationship between pressure, volume, temperature and number of moles of gas of an ideal gas:

$$PV = nRT$$

P = pressure (pascals, Pa)

V = volume (m^3)

n = number of moles of gas (mol)

R = gas constant (8.31 J K^{-1} mol^{-1})

T = temperature (Kelvin, K)

- The ideal gas equation can also be used to calculate the **molar mass** (M) of a gas

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Worked example

Calculate the volume, in dm^3, occupied by 0.781 mol of oxygen at a pressure of 220 kPa and a temperature of 21 °C.

Answer:

- **Step 1:** Rearrange the ideal gas equation to find volume of the gas

$$V = \frac{nRT}{P}$$

- **Step 2:** Convert into the correct units and calculate the volume the oxygen gas occupies

P = 220 kPa = 220 000 Pa

n = 0.781 mol

R = 8.31 J K^{-1} mol^{-1}

T = 21 °C = 294 K

$$V = \frac{0.781\ mol \times 8.31\ J\ K^{-1} mol^{-1} \times 294\ K}{220\ 000\ Pa}$$

= 0.00867 m^3

= **8.67 dm^3**

Examiner Tip

A word about units…

Students often mess up gas calculations by getting their unit conversions wrong, particularly from cm^3 to m^3. Think about what a cubic metre actually is - a cube with sides 1 m or 100 cm long. The volume of this cube is 100 x 100 x 100 = 1 000 000 or **10^6** cm^3

So when you convert from **m^3 to cm^3** you **MULTIPLY by 10^6** and when you convert from **cm^3 to m^3** you **DIVIDE by 10^6** (or multiply by 10^{-6} which is the same thing)

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Worked example

Calculate the pressure of a gas, in kPa, given that 0.20 moles of the gas occupy 10.1 dm^3 at a temperature of 25 °C.

Answer:

- **Step 1:** Rearrange the ideal gas equation to find the pressure of the gas

$$P = \frac{nRT}{V}$$

- **Step 2:** Convert to the correct units and calculate the pressure

  n = 0.20 mol

  $V = 10.1\ dm^3 = 0.0101\ m^3 = 10.1 \times 10^{-3}\ m^3$

  $R = 8.31\ J\ K^{-1}\ mol^{-1}$

  T = 25 °C = 298 K

$$P = \frac{0.20\ mol \times 8.31\ J\ K^{-1} mol^{-1} \times 298\ K}{10.1 \times 10^{-3}\ m^3}$$

  P = 49 037 Pa = **49 kPa** (2 sig figs)

Your notes

Worked example

Calculate the temperature of a gas, in °C, if 0.047 moles of the gas occupy 1.2 dm^3 at a pressure of 100 kPa.

Answer:

- **Step 1:** Rearrange the ideal gas equation to find the temperature of the gas

$$T = \frac{PV}{nR}$$

- **Step 2:** Convert to the correct units and calculate the pressure

n = 0.047 mol

V = 1.2 dm^3 = 0.0012 m^3 = 1.2 x 10^{-3} m^3

R = 8.31 J K^{-1} mol^{-1}

P = 100 kPa = 100 000 Pa

$$T = \frac{100\ 000 \times 1.2 \times 10^{-3}\ m^3}{0.047\ mol \times 8.31\ J\ K^{-1} mol^{-1}}$$

T = 307.24 K = 34.24 °C = **34 °C** (2 sig figs)

Worked example

A flask of volume 1000 cm^3 contains 6.39 g of a gas. The pressure in the flask was 300 kPa and the temperature was 23 °C.Calculate the molar mass of the gas.

Answer:

- **Step 1:** Rearrange the ideal gas equation to find the number of moles of gas

$$n = \frac{pV}{RT}$$

- **Step 2:** Convert to the correct units and calculate the number of moles of gas

P = 300 kPa = 300 000 Pa

V = 1000 cm^3 = 0.001 m^3 = 1.0 x 10^{-3} m^3

R = 8.31 J K^{-1} mol^{-1}

T = 23 °C = 296 K

$$n = \frac{300\ 000\ Pa \times 1 \times 10^{-3}\ m^3}{8.31\ J\ K^{-1} mol^{-1} \times 296\ K}$$

n = 0.12 mol

- **Step 3:** Calculate the molar mass using the number of moles of gas

$$\text{molar mass} = \frac{mass}{moles}$$

$$M = \frac{6.39\ g}{0.12\ mol} = \mathbf{53\ g\ mol^{-1}}\ \text{(2 sig figs)}$$

Examiner Tip

To calculate the temperature in **Kelvin**, add 273 to the Celsius temperature, eg. 100 °C is 373 Kelvin.

Your notes

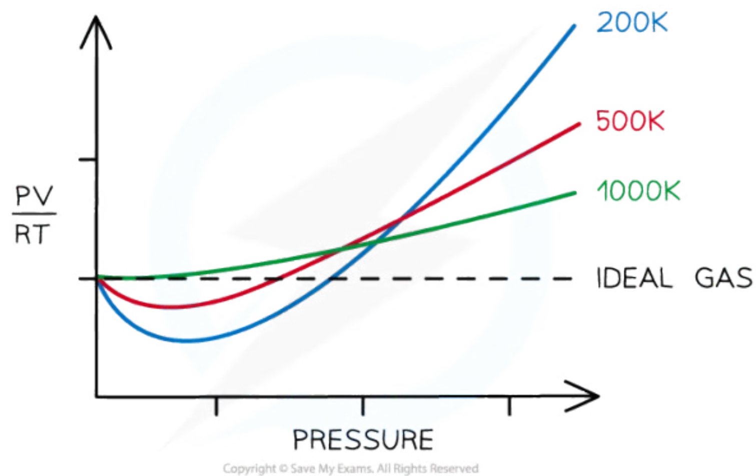
Real Gases

Your notes

Real Gases

- The **ideal gas equation** does not fit all measurements and observations taken at all conditions with real gases
- The relationship between pressure, volume and temperature shows significant deviation from ***PV = nRT*** when the **temperature is very low** or the **pressure is very high**
- This is because the **ideal gas equation** is built on the **kinetic theory of matter**
- The **kinetic theory of matter** makes some key assumptions about the behaviour of gases

Graph to show deviation from Ideal Gas Behaviour

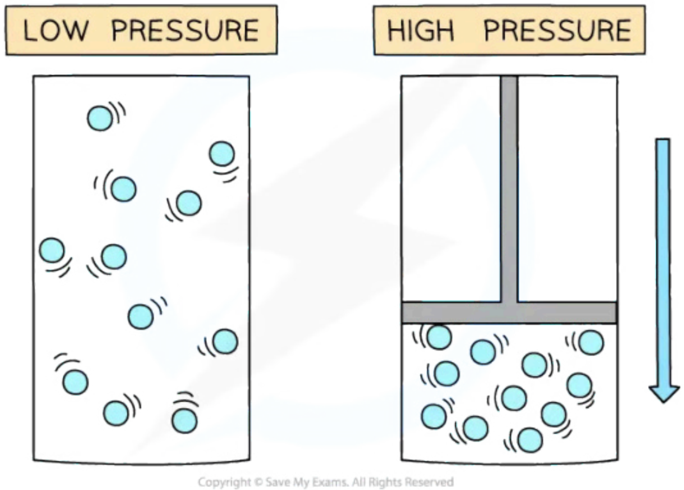


At low temperatures and high pressures real gases deviate significantly from the ideal gas equation. The higher the pressure and the lower the temperature the greater the deviation

Assumptions about volume

- The **kinetic theory** assumes that the volume the actual gas molecules themselves take up is tiny compared to the volume the gas occupies in a container and can be ignored
- This is broadly true for gases at normal conditions, but becomes increasingly inaccurate at low temperatures and high pressures
- At these conditions the gas molecules are very close together, so the **fraction of space** taken up by the molecules is **substantial** compared to the total volume

Diagram to show gas volumes at low and high pressure



At low temperatures and high pressures, the fraction of space taken up by the molecules is substantial

Assumptions about attractive forces

- Another assumption about gases is that when gas molecules are far apart there is very little interaction between the molecules
- As the gas molecules become closer to each other **intermolecular forces** cause **attraction** between molecules
- This reduces the number of collisions with the walls of the container
- The pressure is less than expected by the **ideal gas equation**

Examiner Tip

The ideal gas equation and the gas constant are given in the IB Chemistry Data Booklet.