Cambridge (CIE) A Level Chemistry

Your notes

The Gaseous State: Ideal & Real Gases & pV = nRT

Contents

* Gas Pressure
* Ideal Gases

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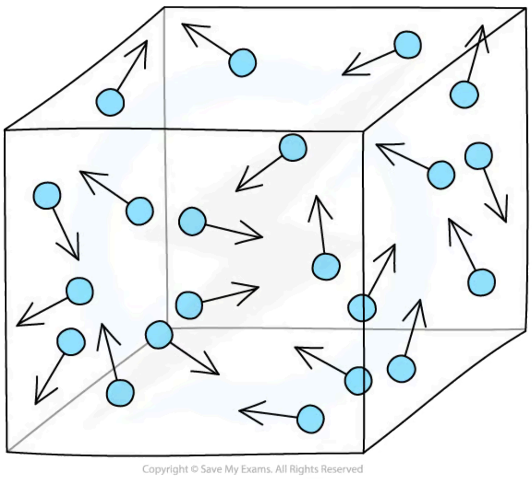
Gas Pressure

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

Illustration of gas pressure

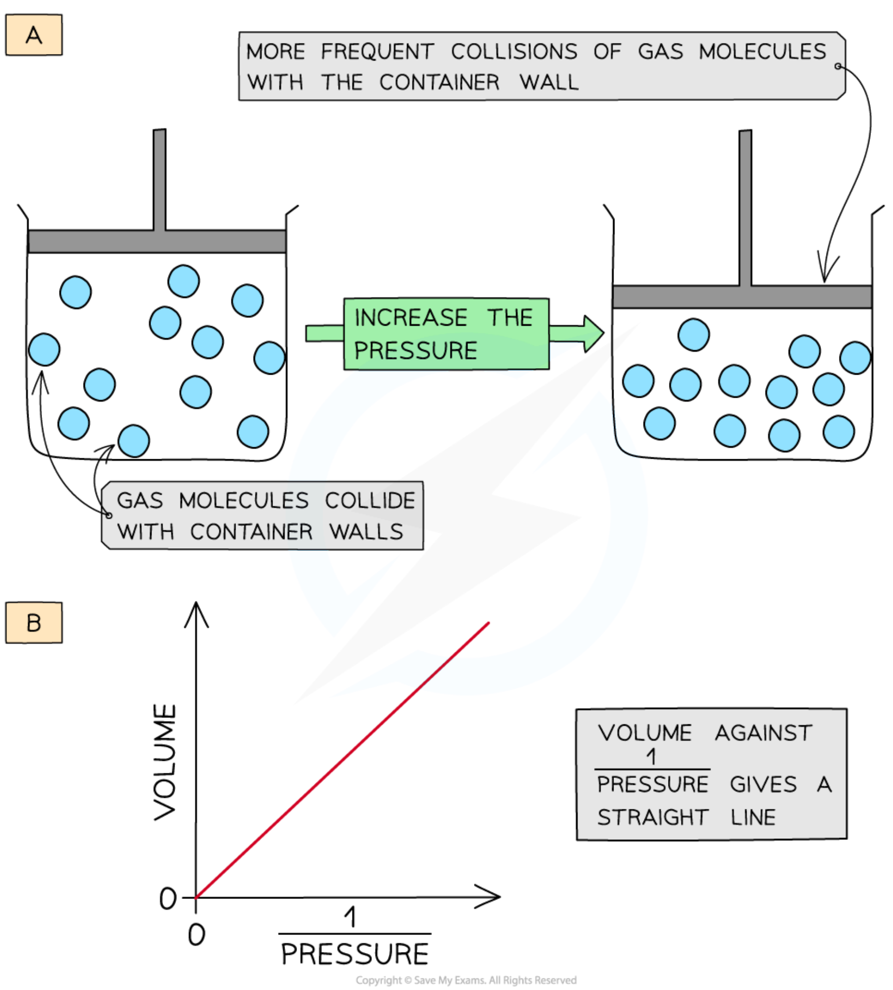


Gas particles exert a pressure by constantly colliding with the walls of the container

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**
- The **volume** is therefore **inversely proportional** to the **pressure** (at constant temperature)
  - A graph of the **volume** of gas plotted against 1/pressure gives a straight line

How decreasing the volume of a gas affects collision frequency

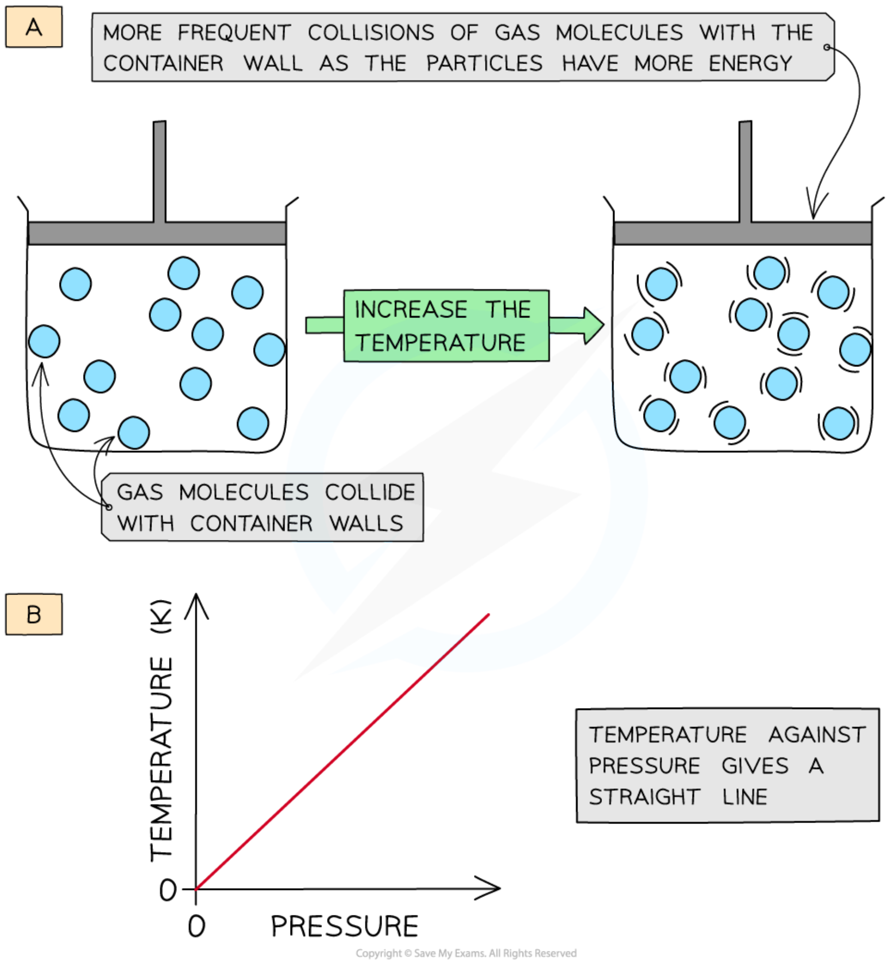


Decreasing the volume of a gas causes an increased collision frequency of the gas particles with the container wall (a); volume is inversely proportional to the pressure (b)

Changing gas temperature

- **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)
  - A graph of the **temperature** of gas plotted against **pressure** gives a straight line

How increasing the temperature of a gas affects collision frequency



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)

Your notes

Ideal Gas Law

Kinetic theory of gases

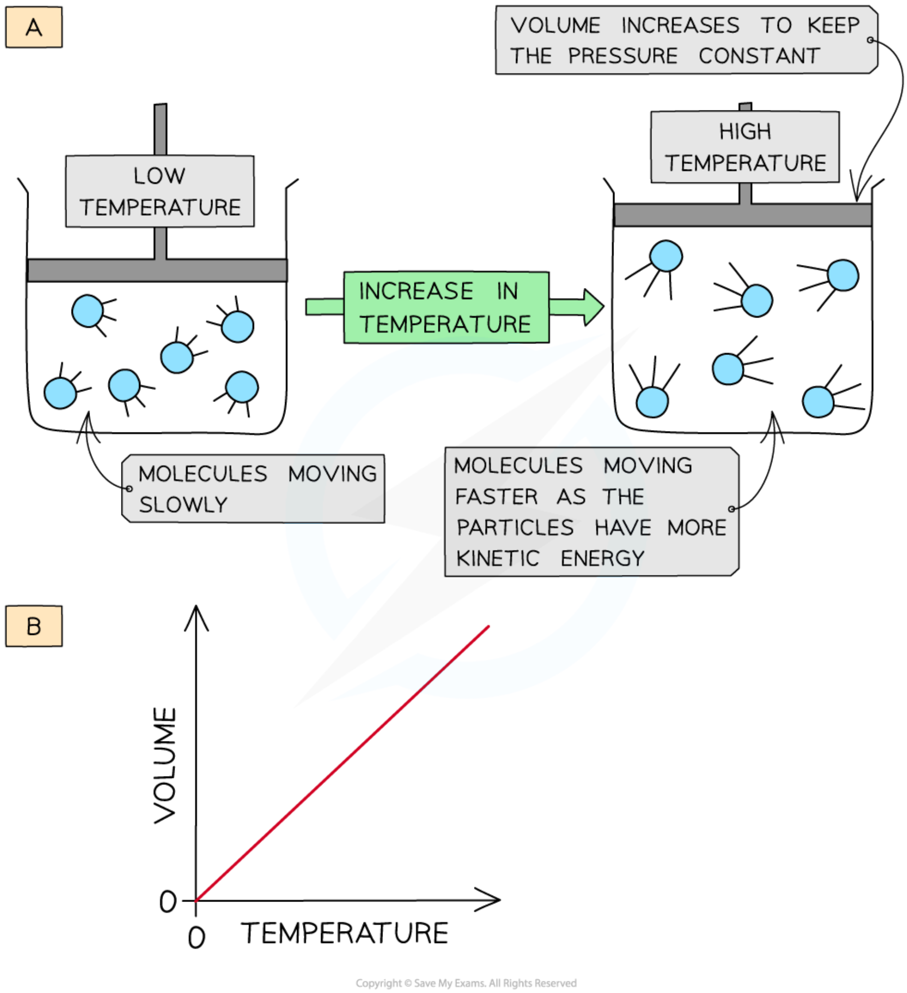
- **The kinetic theory of gases** states that molecules in gases are constantly moving
- The theory makes the following assumptions:
  - The gas molecules are moving very fast and randomly
  - The molecules hardly have any volume
  - The gas molecules do not attract or repel each other (**no intermolecular forces**)
  - No kinetic energy is lost when the gas molecules collide with each other (**elastic collisions**)
  - The temperature of the gas is related 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**

Ideal gases

- The volume that an ideal gas occupies depends on:
  - Its pressure
  - Its temperature
- When a gas is **heated** (at constant pressure) the particles gain more **kinetic energy** and undergo more **frequent collisions** with the container wall
- 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** (at constant pressure)

How the volume of a gas increases upon heating

Your notes



The volume of a gas increases upon heating to keep a constant pressure (a); volume is directly proportional to the temperature (b)

Limitations of the ideal gas law

- At very **high pressures** and **low temperatures** real gases deviate from ideal gas behaviour because:
  - The gas molecules are much closer together.
  - Intermolecular forces (e.g. London forces or dipole-dipole attractions) become significant.
  - The molecules' own volume can no longer be ignored.
- As a result, **two** key assumptions of the kinetic theory break down:
  - There are no intermolecular forces
    - In reality, attractive forces between molecules pull them slightly inward
    - So, they hit the container walls with less force than expected
    - This makes the measured pressure lower than the ideal gas law predicts.

- The volume of the gas molecules can be ignored
  - In reality, the gas particles occupy space
  - So, the actual free volume available for movement is less than predicted by the ideal gas law.

Ideal Gas Equation

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_r) of a gas

Worked Example

Calculating the volume of a gas

Calculate the volume 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 the volume of gas:
  - $V = \dfrac{nRT}{P}$
- **Step 2:** Check that values have the correct units:
  - P = 220 kPa = 220 000 Pa
  - n = 0.781 mol
  - R = 8.31 J K^{-1} mol^{-1}
  - T = 21 °C = 294 K
- **Step 2:** Calculate the volume the oxygen gas occupies:
  - $V = \dfrac{0.781 \text{ mol} \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 294 \text{ K}}{220\ 000 \text{ Pa}}$
  - V = 0.00867 m^3
  - V = 8.67 dm^3

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Worked Example

Calculating the molar mass of a gas

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 relative molecular 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:** Check that values have the correct units:
  - P = 300 kPa = 300 000 Pa
  - V = 1000 cm^3 = 0.001 m^3
  - R = 8.31 J K^{-1} mol^{-1}
  - T = 23 °C = 296 K
- **Step 3:** Calculate the number of moles of gas:
  - $n = \frac{300\ 000\ \text{Pa} \times 0.001\ \text{m}^3}{8.31\ \text{J K}^{-1}\ \text{mol}^{-1} \times 296\ \text{K}}$
  - n = 0.12 mol
- **Step 4:** Calculate the molar mass using the number of moles of gas:
  - $n = \frac{\text{mass}}{\text{molar mass}}$
  - molar mass = $\frac{6.39}{0.12\ \text{mol}}$
  - molar mass = 53.25 g mol^{-1}

Examiner Tips and Tricks

Ideal gases have **zero particle volume** (the particles are really small) and **no intermolecular forces of attraction or repulsion.**

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

Remember: an **ideal gas** will have a **volume** that is **directly proportional** to the **temperature** and **inversely proportional** to the **pressure.**