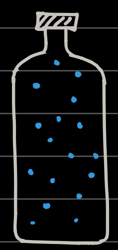
THE GASEOUS STATE

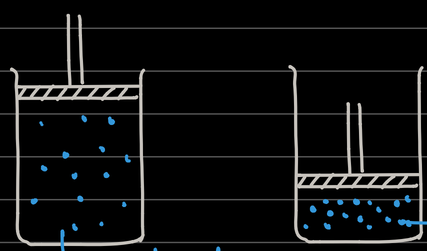
- Ideal gases vs. Real gases
- Calculations using the ideal gas equation

What is a gas? Gases are mostly empty space



1 dm^3 gas → capacity of the container

Actual volume of a gas is the volume the gas would occupy if it was condensed to a liquid.



→ Actual volume

↳ Total volume, basically the volume of the container

Actual volume is < 1% of the total volume

Example :

Steam : 1 mol → 24dm^3
Water : 1 mol → 18g
Density of water = 1 g/cm^3

Vol. of 1 mol water = $\frac{\text{mass}}{\text{density}}$ = $\frac{18g}{1 g/cm^3}$ = 18 cm^3

↱ Actual volume in cm^3

$\frac{18}{24000} \times 100 = 0.075\%$ → Shows that the actual volume is a very small percentage of the total volume and is mostly empty space.
↳ Total volume in cm^3
↳ Thus, it can be compressed.

Standard Temperature & Pressure (STP) : 0°C , 1 atm pressure
Room Temperature & Pressure (RTP) : 25°C , 1 atm pressure

Pressure : is the force that the gas particles exert on a given area of its container

SI Units : Pascals (Pa) or kPa
1 Pa results from a force of 1N acting on an area of 1 square meter.

Atmospheric Pressure: The pressure that the atmosphere exerts on the Earth's surface at sea level.

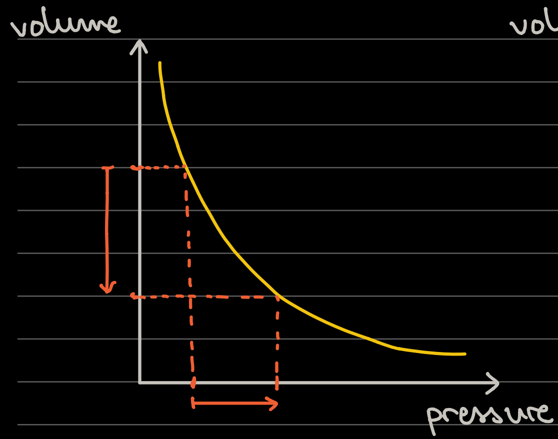
$1 \text{ atm} = 1.01 \times 10^5 \text{ Pa}$

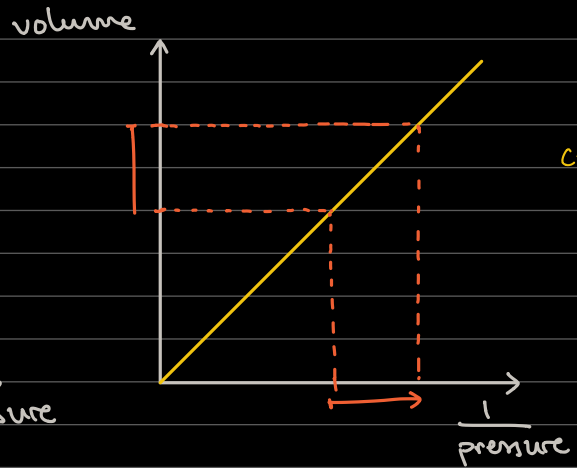
$1 \text{ atm} \approx 1 \times 10^5 \text{ Pa}$

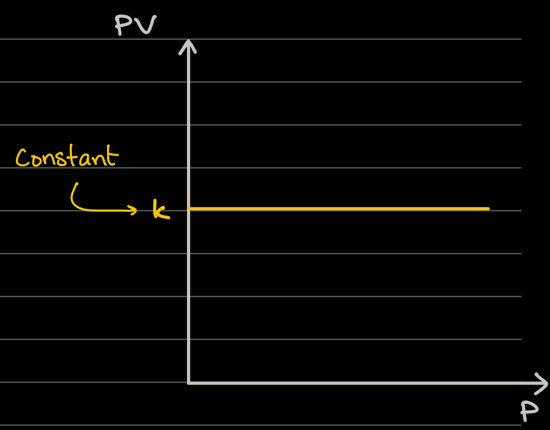
Gas Laws:

Boyle's Law: The volume (v) of a fixed mass of gas at constant temperature is inversely proportional to the pressure on the gas.

$V \propto \frac{1}{P}$ or $VP = \text{constant}$

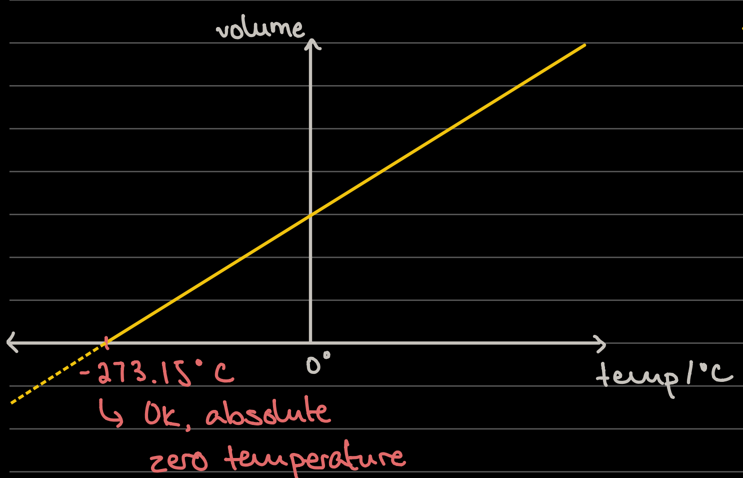






Charle's Law: The volume V of a fixed mass of gas at constant pressure is directly proportional to it's temperature

$V \propto T$ or $\frac{V}{T} = \text{constant}$



- $-273.15°C$ is known as the absolute 0 of temperature
- Below this temperature, Charles' Law predicts that the gas would have a negative volume ($V \propto T$) - but this is not possible
- So we can assume that it is not possible to attain temps below $-273.15°C$.
- So chemists use a temp scale that starts at absolute zero → kelvin
  $K = °C + 273.15$.

Temperature Scales

Absolute (Kelvin)

Celsius Scale (°C)

373 K ⟵ water boils ⟶ 100°C

298 K ⟵ Room temperature ⟶ 25°C

273K ⟵ water freezes ⟶ 0°C

0 ⟵ Absolute Zero ⟶ - 273°C

Avogadro's Law : The volume V of a gas, under identical conditions of constant temperature and pressure, is proportional to the amount in moles "n" of the gas present

$V \propto n$

Putting all three laws together (to come up with the ideal gas equation)

$V \propto \frac{1}{P}$ $V \propto T$ $V \propto n$

Boyle's Charles Avogadro's

$\therefore V \propto \frac{nT}{P}$

Convert it into an equation by adding a proportionality constant called the gas constant R.

$V \propto \frac{nT}{P}$

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

$\boxed{PV = nRT}$ ⟶ Ideal Gas Equation.

↳ An ideal gas follows this equation exactly under all conditions of temperature and pressure.

PV = nRT where....

1. V = volume
   P = Pressure
   n = moles
   T = Temp in Kelvin.
   R = 8.31 J/Kmol ⟶ constant

BUT have to maintain consistency with units when using this equation

If P ⟶ kPa, then V ⟶ dm^3
If P ⟶ Pa, then V ⟶ m^3

Important conversions to remember

$1\ dm^3 = 1000\ cm^3$

$K = {}^\circ C + 273.5$

$1\ atm = 1.01 \times 10^5\ Pa = 101\ kPa$

$1\ m^3 = 1 \times 10^6\ cm^3$

$1\ dm^3 = 1 \times 10^{-3}\ m^3$

$1\ L = 1\ dm^3$

$1\ mL = 1\ cm^3$

Basic Assumptions of the behaviour of an ideal gas using the kinetic theory of gases

Important

1. The gas particles or molecules have negligible size or volume compared to the volume of the gas

Important

2. The particles have no intermolecular forces of attraction between them

3. The collisions between the molecules are perfectly elastic, ie. there is no loss of K.E. when the molecules collide with each other and with the walls of the container

4. The gas particles are in constant random motion.

5. The pressure of a gas is due entirely to collisions between the molecules and with the walls of the container

Real Gases vs. Ideal Gases

A real gas deviates from ideal gas behaviour because 2 of the basic assumptions of the kinetic molecular theory are not entirely true

1. In a real gas, the molecules have a certain volume, and hence they have a significant size

2. There are attractive forces between the particles, though they are usually weak

When is a real gas most like an ideal gas?

a) At low pressure and,
b) at high temperatures

a) Low pressure ($V \propto \frac{1}{P}$)

- At low pressure, the gas molecules are far apart so they have a negligible size / volume since the total volume is very large.

b) High Temp

- particles have a large amount of kinetic energy
- they will move faster
- when they collide, they won't "stick together"
- Intermolecular forces would be insignificant as particles will have sufficient kinetic energy to overcome these intermolecular forces.

When does a real gas deviate most from an ideal gas?

a) At high pressure, and

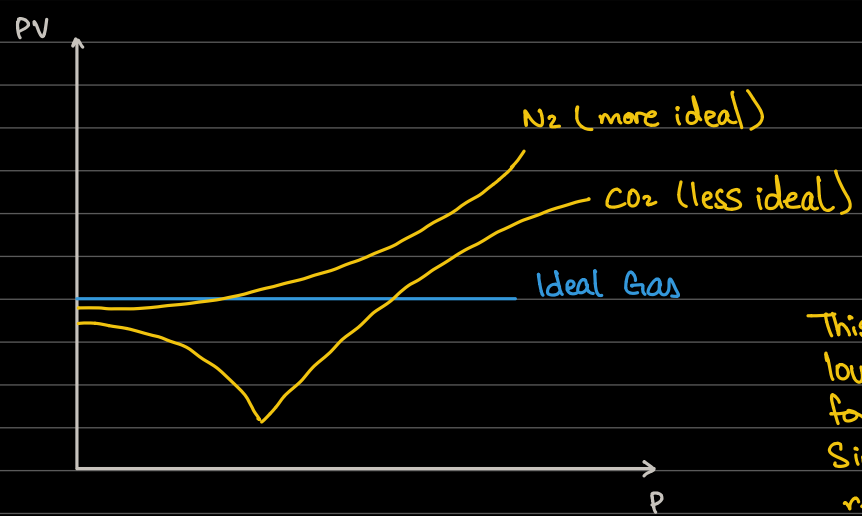
b) at low temperature

a) High Pressure (negates the 1st assumption)

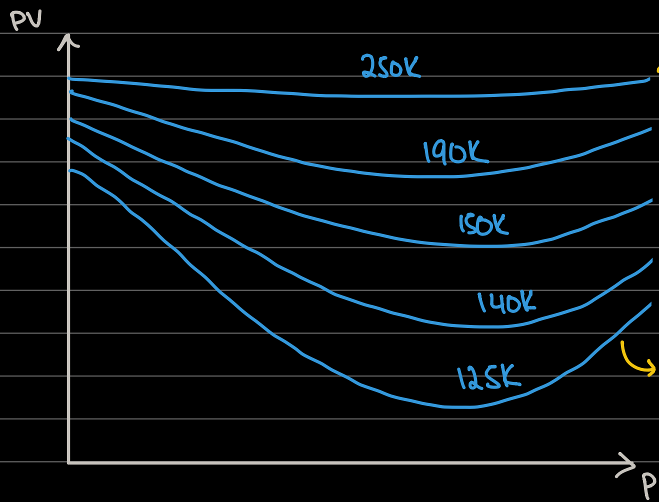
- At high pressure (less volume), the gas particles / molecules are packed closer together, so the actual volume of the gas becomes increasingly significant in proportion to the overall volume of the gas.

b) Low temperature (negates the 2nd assumption)

- particles have less kinetic energy so they tend to "stick together" on colliding with each.
- ie. intermolecular forces become more significant (nearer to it's boiling point)



This is because N_2 has a lower Mr, hence weaker ID-ID forces.
Since intermolecular forces are relatively less significant, N_2 behaves more like an ideal gas than CO_2.



→ most like an ideal gas at higher temperatures
↳ particles do not "stick together"
IMF become insignificant

↳ least like an ideal gas
(N_2 deviates most from an ideal gas at low temperatures)
↳ Particles have less KE, they tend to stick together upon collision.
↳ IMF become significant

Example:

List the following in decreasing order of similarity to an ideal gas

CH_4, H_2, He, CO_2, NH_3

Most like ideal gas → the one with the weakest IMF = Helium
why? Because He is a monoatomic gas, thus less volume than Hydrogen

Next → Hydrogen
3rd → CH_4 why? Lower Mr, so weaker ID-ID forces
4th → CO_2 why? Relatively higher Mr, so relatively stronger IMF
5th → NH_4 why? Has the strongest IMF forces: Hydrogen bonds