

2023 JC1 H2 CHEMISTRY (9729) CORE IDEA 2: STRUCTURE AND PROPERTIES Topic 4: THE GASEOUS STATE

Name: _____

Civics Group: _____

Students should be able to:

- (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
- (d) use Dalton's Law to determine the partial pressures of gases in a mixture

Lecture Content

This topic is not included in the H1 Chemistry Syllabus.

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REFERENCES

- 1 Chemistry (for CIE AS & A Level) by Peter Cann & Peter Hughes
- 2 Chemistry³: Introducing Inorganic, Organic and Physical Chemistry by Andrew Burrows, John Holman, Andrew Parsons, Gwen Pilling & Gareth Price
- 3 A-Level Chemistry by E.N. Ramsden

1 Introduction

Gases tend to have the following physical properties:

- Gases have much lower density than solids or liquids. (Why is density low for gases?)
- Gases assume the volumes and shapes of their containers. (Why ONLY gases assume the volumes of their containers? Why liquids do not assume the volume leh?)
- Gases are highly compressible and their volumes change greatly when a pressure is applied. (Why are gases highly compressible?)
- Gases exert pressure equally in all directions. (Why gases exert pressure equally in all directions?)



Note: <u>ALL the above is actually because gases have MUCH weaker (as compared to liquid</u> <u>ah) intermolecular forces, which can be hydrogen bonding, permanent dipolepermanent dipole or instantaneous dipole-induced dipole interactions</u>!!!

Making Thinking Visible

- Q: Do you know what is pressure?
- A: Pressure refers to the <u>force that the particles exert on a surface per unit area</u>. Hence, pressure measured is not due to the collision between the particles.

Q: Since pressure refers to the force that the particles exert on the surface per unit area, does that mean that if there are more particles hitting per unit area, the pressure will be higher?

A: Yes!!! That is why the measured pressure is directly proportional to the amount of particles, in moles (n).

In addition, **pressure is directly proportional to temperature (T)** because the higher the temperature, the greater the kinetic energy of the particles (KE_{average} α T). The faster the particles move, the "harder" it hits against the surface. The particles also strike the surface more frequently.

Lastly, the **pressure is inversely proportional to the volume of the container (V)** because the greater the volume, the further the particles are spread out, the lower the frequency of collisions as the particles will take a "longer time" to hit the surface again.

- **Q:** Sir, if *p* of a gas is affected by *n*, *V* and *T*, does it mean that *V* of a gas is being affected by *p*, *n* and *T*?
- A: Certainly Eunoians, we will later discuss the interconnectedness of p, V, n and T in the ideal gas equation, pV = nRT.

S.I. Units and Common Units of Measurement for the Properties of Gases

The physical behaviour of a sample of gas can generally be described completely by four variables:

The Gaseous State

(a) Pressure exerted by the gas, p				
	S.I. Unit	pascal	<mark>(Pa)</mark>	1 Pa = 1 <mark>N m⁻²</mark>
	Other commonly used units	atmosphere	(atm)	1 atm = 101 325 Pa
		bar	(bar)	<mark>1 bar = 10⁵ Pa</mark>
		torr	(torr or mm Hg)	760 torr = 760 mm Hg = 1 atm = 101 325 Pa

(b) Volume occupied by the gas, V

S.I. Unit	cubic metre	(<mark>m³</mark>)	
	cubic decimetre	(dm³)	<mark>1 dm³ = 10⁻³ m³</mark>
Other	cubic centimetre	(cm ³)	<mark>1 cm³ = 10⁻⁶ m³</mark>
used units	litre	(L)	1 L = <mark>1 dm³ = 10⁻³ m³</mark>
	millilitre	(mL)	$1 \text{ mL} = 1 \text{ cm}^3 = 10^{-6} \text{ m}^3$

- (c) Amount of the gas, *n* S.I. Unit mole (mol)
- (d) Temperature of the gas, *T*

S.I. Unit	kelvin	(<mark>K</mark>)			
Other commonly used units	degree Celsius	(°C)	<u><i>T</i></u> / ^o C + 273 = <i>T</i> / K		

Self-Check 1A					
Express the follo	owing quantities in	their S.I. units.			
(a) 2 atm	(b) 500 mL	(c) 22.7 dm ³	(d) 30 cm ³	(e)	25 °C
(a) 2 x 10⁵ Pa	(b) 500x10 ⁻⁶ m ³	(c) 22.7x10 ⁻³	m³ (d) 30x10 ^{-/}	⁶ m ³	(e) 298 K

2 The Gas Laws

LO (c) state and use the general gas equation pV = nRT in calculations, including the determination of M_r

Some of the first quantitative investigations were conducted in the $17^{th} - 18^{th}$ century to study the behaviour of gases. The relationship between the different macroscopic properties of gases (volume, temperature and pressure) are expressed via the different gas laws.

2.1 Boyle's Law

The volume V of a fixed mass of gas at a constant temperature is inversely proportional to the pressure p of the gas.



In other words, for a fixed mass (or amount) of gas at constant temperature, when its volume increases, its pressure decreases, and vice versa.





A balloon occupies a volume of 600 cm³. Assuming that temperature remains constant, what volume will the balloon occupy if the surrounding pressure is reduced so that the pressure in it is reduced to one-third of its starting value?

Let the initial pressure be *p*.
By Boyle's law,
$$p_1V_1 = p_2V_2$$

 $p(600) = \left(\frac{1}{3}p\right)(V_2)$
 $V_2 = 3 \times 600$
 $= \frac{1800 \text{ cm}^3}{2}$

2.2 Charles's Law

The volume *V* of a **fixed mass of ga**s at a **constant pressure** is **directly proportional** to its temperature *T* (in kelvin).

 $V \propto T$ $\frac{V}{T} = k$ where k is a constant

- In other words, for a fixed mass (or amount) of gas at constant pressure, when its volume increases, its temperature increases, and vice versa
- * For a given mass of gas under two different sets of conditions at constant pressure,

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

Graphically, Charles's Law can be represented as such:



At constant *p* and n,

V = kT(Resembles y = mx graph)

- The absolute temperature scale or Kelvin scale was defined based on the linear relationship between gas volume and temperature. The Kelvin scale starts at absolute zero (i.e. $0 \text{ K} = -273 \text{ }^{\circ}\text{C}$). Each degree in this scale has the same size as a degree on the Celsius scale.
- To convert temperature from degree Celsius to kelvin,

T/K = T/°C + 273

When gas volume is plotted against temperature in degree Celsius, the graph is as follows:



Note: The relationship $V \propto T$ is only applicable when temperature is expressed in kelvins (NOT degree Celsius).

Example 2B

32.0 cm³ of carbon dioxide from a chemical reaction is collected in a syringe at 0 °C and 1.00 atm. What would be its volume at r.t.p (20°C and 1 atm)?

At r.t.p., T = 20 °C, p = 1 atm By Charles's law,

 $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\frac{32.0}{0+273} = \frac{V_2}{20+273}$ $V_2 = \frac{34.3 \text{ cm}^3}{20}$

2.3 Avogadro's Law

Under the same conditions of temperature and pressure, equal volumes of all gases will contain equal numbers of molecules.

 $V \propto n$ where *n* is the amount (in moles) of gas present

- * For example, the number of O_2 molecules in 1 dm³ of O_2 gas = the number of CO_2 molecules in 1 dm³ of CO_2 gas.
- * At constant temperature and pressure,

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Graphically, Avogadro's Law can be represented as:



At constant *T* and *p*,

V = kn(Resembles y = mx graph)

2.4 The Ideal Gas Equation

Sombining Boyle's law, Charles's law and Avogadro's law,

 $V \propto \frac{nT}{p}$ therefore, pV = nRT where R is the molar gas constant $(R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1})$

An ideal gas is one whose behaviour follows the ideal gas equation exactly under any pressure and temperature.

Important:	When using the ideal gas equation, each quantity needs to be expressed in its
1 - 1	S.I. unit.

From $pV = nRT$,	if n and T constant	$\Rightarrow p \alpha \frac{1}{v}$	Boyle's Law
	If n and p constant	$\Rightarrow V \alpha \dot{T}$	Charles's Law
	If p and T constant	\Rightarrow V α n	Avogadro's Law
	If n and V constant	\Rightarrow p α T	
	If p,V and T constant	$\Rightarrow \frac{\mathrm{pV}}{\mathrm{T}} = \mathrm{nR}$	
	If p,V, T and n constant	$\Rightarrow \frac{\mathrm{pV}}{\mathrm{nT}} = R$	

Example 2C

Calculate the molar volume of an ideal gas at 0 °C and 1 bar (i.e. under standard temperature and pressure, s.t.p.).

Converting the quantities to S.I. units:

Pressure	1 bar = 10 ⁵ Pa	
Volume	?	
Temperature	0 °C = 0 + 273 = 273 K	

pV = nRT $pV_{m} = RT$ $(10^{5})(V_{m}) = (8.31)(273)$ $V_{m} = 0.0227 \text{ m}^{3} \text{ mol}^{-1} = 22.7 \text{ dm}^{3} \text{ mol}^{-1}$

Example 2D

A sample of gas was heated from 25 °C until both its pressure and its volume doubled. What was the new temperature?

Since the amount of the gas remains constant, $pV = nRT \Rightarrow \frac{pV}{T} = nR = \text{constant}$

Let the initial pressure and volume be p_1 and V_1 ,

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$
$$\frac{p_1 V_1}{(25+273)} = \frac{(2p_1)(2V_1)}{T_2}$$
$$T_2 = 4(25+273)$$
$$= 1192 \text{ K} = \underline{1190 \text{ K}} \quad (3 \text{ s.f.})$$

Example 2E

The volume occupied by 0.10 g of a gas is found to be 83.1 cm³ (measured at 1.0×10^5 Pa and 27 °C). Assuming that the gas behaves ideally, what is the relative molecular mass of this gas?

$$pV = nRT$$

$$pV = \frac{\text{mass of gas}}{\text{molar mass of gas}} \times RT = \frac{m}{M}RT$$

$$(1.0 \times 10^5)(83.1 \times 10^{-6}) = \frac{0.10}{M}(8.31)(27 + 273)$$

$$M = \frac{(0.10)(8.31)(27 + 273)}{(1.0 \times 10^5)(83.1 \times 10^{-6})}$$

$$= 30.0 \text{ g mol}^{-1}$$

$$M_r \text{ of gas} = \underline{30.0} \text{ (3 s.f.)}$$

Example 2F

Determine the density of H_2S at 20 °C and 2.00 atm, assuming that it behaves like an ideal gas under the stated conditions.

$$pV = nRT$$

$$pV = \frac{\text{mass of gas}}{\text{molar mass of gas}} \times RT = \frac{m}{M}RT$$

$$pM = \frac{m}{V}RT$$
density, $\rho = \frac{m}{V} = \frac{pM}{RT}$

$$= \frac{(2.00 \times 101325)(1.0 \times 2 + 32.1)}{(8.31)(20 + 273)}$$

$$= \frac{2.84 \times 10^3 \text{ g m}^{-3}}{2.84 \times 10^3 \text{ g m}^{-3}}$$

Self-Check 2B

- When an evacuated glass bulb of volume 63.8 cm³ was filled with a gas at 24 °C and 99.5 kPa, its mass was found to increase by 0.103 g. Is the gas likely to be ammonia, nitrogen or argon? [Argon]
- 1 cm³ of ice was heated to 323 °C at a pressure of 1 atm. Given that the density of ice is 1.00 g cm⁻³, what was the volume of steam produced? [2.72 dm³]
- 3. A fluoride **X** is formed when krypton reacts with fluorine. Given that 0.100 g of **X** occupies 104 cm³ at 2.00 × 10⁴ Pa and 305 K, what is the relative molecular mass and molecular formula of **X**? [121.9; KrF₂]

2.5 Dalton's Law of Partial Pressures

- LO (d) use Dalton's Law to determine the partial pressures of gases in a mixture (see also Section 9)
 - In a mixture of gases that do not react with one another, each gas behaves as if it was the only gas present.



Gas A only pressure = p_A



Gas B only pressure = $p_{\rm B}$



Mixture of **Gases A and B** pressure = $p_A + p_B$

each gas in the mixture exerts the same pressure as if it were the only substance in the container In a mixture of gases which <u>do not react chemically</u>, the total pressure is the sum of the partial pressures of the components.

$$\mathcal{D}_{\text{total}} = \mathcal{P}_{\text{A}} + \mathcal{P}_{\text{B}} + \mathcal{P}_{\text{C}} + \dots + \mathcal{P}_{\text{N}} = \sum_{i=1}^{N} \mathcal{P}_{i}$$

where p_i is the partial pressure of any component, *i*, in the mixture

- The partial pressure is the pressure that would be exerted if the gas was alone in the container. We can determine the partial pressure of each component gas using its mole ratio. (Partial pressure = part of the total pressure)
- * Assuming a non-reacting mixture of gas **A** and gas **B** contained in a volume, *V*, at temperature, *T*,



Mole Fraction and Partial Pressure

* The **mole fraction** (fraction calculated using moles) of a component A, χ_A , in a mixture is the ratio of the number of moles of A to the total number of moles of all components present in the mixture. Mathematically, this is shown as:

amount of component A	<i>n</i>
λ_{A}^{-} total amount of all components in mixture	$n_{\rm total}$

* If there is a non-reacting mixture of gas A and gas B in a container, then:

$$\frac{p_{A}}{p_{\text{total}}} = \frac{n_{A}RT}{V} \div \frac{(n_{A} + n_{B})RT}{V}$$
$$= \frac{n_{A}}{n_{A} + n_{B}}$$
$$= \chi_{A} \text{ (where } \chi_{A} \text{ is the mole fraction of gas A)}$$

* The partial pressure of gas A is thus:

$$p_A = \chi_A \times p_{total}$$

Example 2G

A mixture of 0.5 mol of helium, 2.0 mol neon and 2.5 mol of argon has a total pressure of 100 kPa. What is the partial pressure of each gas in the mixture?

$$p_{\text{He}} = \frac{0.5}{0.5 + 2.0 + 2.5} \times 100 = 10 \text{ kPa}$$
$$p_{\text{Ne}} = \frac{2.0}{0.5 + 2.0 + 2.5} \times 100 = 40 \text{ kPa}$$
$$p_{\text{Ar}} = \frac{2.5}{0.5 + 2.0 + 2.5} \times 100 = 50 \text{ kPa}$$

Example 2H

Flask X contains 1 dm³ of helium at 2 kPa pressure and Flask Y contains 2 dm³ of neon at 1 kPa pressure. If the flasks are connected at constant temperature, what is the final pressure?

Method 2

Method 1

Total no. of moles do not change. Final pressure = $p_{He} + p_{Ne}$ Before connection, $n_{\text{He}} = \frac{pV}{RT} = \frac{(2)(1)}{RT} = \frac{2}{RT}$ Since temperature is constant: For helium, $p_i V_i = p_f V_f$ $n_{\rm Ne} = \frac{pV}{RT} = \frac{(1)(2)}{RT} = \frac{2}{RT}$ $1 \times 2 = (1+2)p_{f}$ After connection, $n_{\text{total}} = n_{\text{He}} + n_{\text{Ne}}$ For neon, $p_i V_i = p_f V_f$ $\frac{(p_{\text{final}})(V_{\text{final}})}{RT} = \frac{2}{RT} + \frac{2}{RT}$ $2 \times 1 = (1+2)p_{f}$ $p_{\text{final}} = \frac{2+2}{1+2}$ Final total pressure = 0.6667 + 0.6667= 1.33 kPa

W-Making Thinking Visible

Q: Are the above two methods different in working principles?

A: Nope! They are the same. Let me show you: From method 1, after connection,

 $n_{\rm total} = n_{\rm He} + n_{\rm Ne}$ $\frac{p_{\text{final}}V_{\text{final}}}{\frac{\text{RT}}{\text{RT}}} = \frac{p_{\text{He,initial}}V_1}{\frac{\text{RT}}{\text{RT}}} + \frac{p_{\text{Ne,initial}}V_2}{\frac{\text{RT}}{\text{RT}}}$ $p_{\text{final}}V_{\text{final}} = p_{\text{He, initial}}V_{\text{He, initial}} + p_{\text{Ne, initial initial}}V_2$ (since T is constant) $p_{\text{final}}(1+2) = 2x1 + 1x2$ (in Method 2)

 $p_{\rm f} = 0.667 \, \rm kPa$

 $p_{\rm f} = 0.667 \, \rm kPa$

= 1.33 kPa

Method 1 is preferred as most of the time in A-level, the question set will have a constant amount of gases (in moles), before and after the change.

Important concepts to remember:

- (i) I need to know S.I. units for pressure (Pa), volume (m³), temperature (K) and amount (in moles)
- (ii) I need to know how to convert for eg. from °C to K, from cm³ to m³, from atm to Pa, etc.
- (iii) I need to know how to sketch various types of graphs by modifying pV = nRT and know that $n = \frac{mass}{molar mass}$ (where molar mass is numerically the same as M_r).
- (iv) I need to use pV = nRT for various types of calculation and in the process be mindful of units conversion.
- (v) Dalton's Law of Partial Pressure is a very important concept to remember for future topics.

3 The Kinetic Molecular Theory of Gases

LO (a) state the basic assumptions of the kinetic theory as applied to an ideal gas.

The ideal gas equation was obtained using an *empirical* approach by combining the experimentally determined gas laws (Boyle's Law, Charles's Law, Avogadro's Law). While it enables us to determine what happens to a gas when conditions change, it does not tell us what is happening at the molecular level, and why.

In the 19th century, a group of scientists including Ludwig Boltzmann and James Clerk Maxwell developed the **kinetic molecular theory of gases** to try to explain gas behaviours as described by the gas laws and the ideal gas equation. This theory is based on a model that assumes that:

- 1. The **gas particles have no volume**, *i.e.* they are point masses.
- 2. There are no forces of attraction between gas particles.
- 3. When the particles collide, the **collision is perfectly elastic**. i.e., the particles bounce apart after they collide with **no loss of kinetic energy**.
- 4. The gas particles are in **constant random** motion.

As a result, the <mark>average kinetic energy of the particles is proportional to the temperature (in</mark> <mark>Kelvin).</mark>

A Maxwell-Boltzmann distribution curve shows how energy is distributed among the particles in a system. The total area under a Maxwell-Boltzmann distribution curve is proportional to the total number of particles present. At a given temperature, the particles have a range of kinetic energies.

When temperature increases from 298 K to 500 K, the particles have a larger range of kinetic energy and the average kinetic energy of the gaseous particles increases. As such, the Maxwell-Boltzmann distribution curve at the higher temperature broadens and flattens out.



Making Thinking Visible

- **Q:** Why must the peak of the curve be lower at higher temperature?
- A: Now, the y-axis represents the number of particles or fraction of particles with a particular energy, hence, the area under the curve corresponds to the total number of particles or is equal to 1. Imagine each curve is approximated to a triangle. So, if the temperature is increased without changing the number of particles, in order to have a broader curve (a triangle with a greater base), the peak of the curve (height of a triangle) must be lower.
- **Q:** By the way, is the fraction here referring to mole fraction?
- A: Yup!

4 The Behaviour of Real Gases

- LO (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
 - For an ideal gas, $\frac{pV}{nRT}$ will have a constant value of 1 (since pV = nRT).
 - An ideal gas (one which obeys the ideal gas equation under all conditions) is hypothetical in nature. All real gases do not obey the ideal gas equation. Deviations from the ideal gas behaviour can be observed by plotting $\frac{pV}{nRT}$ versus *p*.



- Deviation from ideality is due to:
 - 1) The intermolecular forces of attraction between gas particles.
 - 2) The volume of each gas particle relative to total volume occupied by the gas becoming more significant.

(Recall assumptions of kinetic molecular theory of gases outlined in Section 3)

Real gases deviate most from ideal gas behaviour at high pressure and low temperature.



-Making Thinking Visible

- **Q:** Is the space for real gas particles to move equal to the volume of container?
- A: No! Ideal gas particles have no volume but real gas particles have. Hence, it is IMPORTANT to know the difference between volume of gas particles and volume of gas. The latter refer to the volume of container that the gas occupied!

4.1 Effect of Low Pressure (Real gas becomes more ideal)



gas particles are far apart, the intermolecular forces is weaker and the volume of gas particles is insignificant compared the volume of container.

<u>Making Thinking Visible</u>

Q: At the same temperature, same pressure and for the same amount of gases, can you show us why the volume of a real gas is larger than the volume of an ideal gas?

A: Take note of the following:

Vol. of gas = Vol. of gas particles + Vol. of empty spaces between the gas particles

The volume of empty spaces between the gas particles is the **same** for both real and ideal gases. But the volume of ideal gas particles is **zero**, hence volume of a real gas is larger than volume of an ideal gas.



(@)-<u>Making Thinking Visible</u>

- **Q:** Hmmm....so before the pressure becomes too high, we only consider the intermolecular forces as this is the main factor that causes the gas to become less ideal. But as the pressure keeps on increasing, the volume of gas decreases, causing the volume of the gas particles to be more significant as compared to the volume of empty spaces. So, is the volume of gas particles the main factor that causes the deviation from ideality?
- A: Absolutely! Under most circumstances (at low pressure), the more important factor to take note is actually intermolecular forces which has been covered in Chemical Bonding. We will see later that the stronger the intermolecular forces, the greater the deviation from ideal behaviour. Your insight also explains why after liquefication, it is more difficult to compress liquid than gas.
- **Q:** Can I plot $\frac{pV}{nRT}$ against p for the above plot?
- A: Yes, you can! But take note that for ideal gas the $\frac{pV}{nRT}$ value equals 1.

4.2 Effect of High Temperature (Real gas becomes more ideal)

For a constant amount of a real gas, in this case, CO₂, at varying temperatures:



At high temperatures, <u>kinetic energy of the gas particles increases</u> and the particles move faster. This causes the <u>intermolecular forces of attraction between the particles</u> to be insignificant (assumption 2 becomes valid). The number of collision and force exerted on wall of container will increase as compared to at lower temperature, *pV* will approach that of an ideal gas. Hence at <u>higher temperatures</u>, there will be <u>lesser</u> deviation from ideality.

4.3 Effect of Strong Intermolecular Forces

Wunder the same conditions (same amount and same temperature), different gases deviate from ideal gas behaviour to different extents due to **differences in the strengths of the intermolecular forces of attraction**.



- Strong hydrogen bonds exist between H₂O molecules while weak instantaneous dipole-induced dipole interactions exist between H₂ and CO₂ molecules.
- The hydrogen bonds between H₂O molecules are stronger as compared to the instantaneous dipole-induced dipole interactions between H₂ and CO₂ molecules (assumption 2 less valid for H₂O). Hence, H₂O deviates more from ideality compared to H₂ and CO₂ at low pressures.
- Since CO₂ has more electrons than H₂, the strength of the instantaneous dipole-induced dipole interactions between CO₂ are stronger than between H₂. Hence, CO₂ deviates more from ideality compared to H₂ at **low pressures**. As the instantaneous dipole-induced dipole interactions between H₂ molecules is too weak, the pV values do not deviate below that for the ideal gas.

See <u>Appendix 3</u> for further discussion on deviations from the ideal gas laws.

Making Thinking Visible

- **Q:** Is it important to know which gas molecule has a larger molecular volume than another gas molecule at higher pressures?
- A: No! At A-level the most important factor to consider will be the intermolecular forces as the deviation from ideality can tell us more about the different strengths of the IMF. Take for instance, if you compare the H₂ graph and the CO₂ graph, before the cross-over, we cannot be sure whether the molecular volume of H₂ is larger than the molecular volume of CO₂. Then after the cross-over, we also cannot be sure that the molecular volume of CO₂ is larger than the molecular volume of H₂.

Example 4A

Rank the following gases in order of decreasing ideality, with the most ideal first. Explain the reasons for your order.

 CH_4 HF HCl

- CH₄ is non-polar with weak instantaneous dipole- induced dipole interactions between its molecules.
- HCl is polar with strong permanent dipole-permanent dipole interactions between its molecules.
- Compared to HCl molecules, there is stronger hydrogen bonding between HF molecules (F is highly electronegative).
- Iterative Hence, in terms of ideality, CH₄ is most ideal (deviates least from ideality), followed by HCl and HF.

Important concepts to remember:

- (i) I need to mug AT LEAST three assumptions of the Kinetic Molecular Theory of Gases.
- (ii) I need to know how to sketch the Maxwell-Boltzmann distribution curve as this is important for Chemical Kinetics.
- (iii) I need to know the two main factors that cause gases to be non-ideal:
 - There is intermolecular forces between the gas particles.
 - The volume of <u>gas particles</u> is significant.
- (iv) I need to know the two conditions that make gases to be more ideal:
 - Low pressure: Gas particles are further apart, intermolecular forces are insignificant and volume of particles are insignificant compared to volume of container. <u>The lower the pressure, the smaller the deviation from ideal behaviour</u>
 - High temperature: Gas particles have higher kinetic energy, intermolecular forces are insignificant. <u>The higher the temperature, the smaller the deviation from ideal behaviour</u>.
- (v) I need to know the stronger the intermolecular forces, the greater the deviation from ideal behaviour.

5 Real Life Applications

Why is it important to know how gases behave under different conditions?

Find out the answers to the questions below to understand how important the gas laws are!



Why must scuba divers return to the surface very slowly at the end of their dive?

Reason:

The pressure at the surface is lower than below sea-level. Hence, the solubility of gases in the blood is lower. Therefore, if the surfacing is too fast, the dissolved gases will form bubbles in the blood. It can be fatal.

Why must trucks transporting propane (a highly flammable gas) be subjected to weight tests before they travel during the summer months?

Reason:

Higher mass means more gas particles (more moles). At higher temperature, with greater n, the pressure will increase since volume of container does not change (p α nT).





Why might it not be a good idea to bring a bag of chips onboard an aeroplane?

Reason:

At higher height above sea-level, atmospheric pressure is lower, hence, the higher internal pressure inside the bag may cause the bag to rupture apart.

Why do we find it harder to breathe at high altitudes?

Reason:

At high altitudes, the atmospheric pressre is lower. In order for the gas to go into the lung, the pressure must be high enough. So, at a lower atmospheric pressure, it is difficult for gas particles to go into the lung.





Why does water boil at a lower temperature at higher altitudes?

Reason:

During boiling, the vapour pressure of the liquid must be equal to the atmospheric pressure. At high altitudes, the atmospheric pressre is lower. Hence, a lower temperature is needed for the water to have a vapour pressure equal to the atmospheric pressure for boiling to take place (p α T).

APPENDIX

Scientific Law vs. Theory

What are they? What is the difference?

In our everyday usage, the word "theory" sometimes means a guess or a hunch, and implies an explanation which is uncertain, while the word "law" is often used to refer to things that are immutable truths. In Science, these two terms have very specific meanings:

	What does it do?	Examples
Scientific Law	Describes or predicts what will happen based on a set of given conditions	Laws of motion (Isaac Newton), Law of universal gravitation (Isaac Newton), Law of conservation of mass (Antoine Lavoisier), Law of definite proportions (Joseph Proust)
Scientific Theory	Explains <u>why</u> things happen	Atomic theory, Big Bang theory, climate change, evolution by natural selection, plate tectonics

Changes to Laws & Theories in Science

Scientific knowledge is reliable and durable, but is subjected to revision in the light of new evidence from technological advancements. Scientific laws may be revised when new and unexpected observations are made. Scientific theories may also undergo changes over time – remember how the idea of the atom evolved from the time of the ancient Greeks? There can be multiple theories developed to explain a phenomenon, but only the one(s) that can explain most of the observations made will be favoured by scientists.

To find out more about the relationship between scientific laws and theories, check out:

What's the difference between a scientific law and theory? (Video)



<u>https://ed.ted.com/lessons/what-s-the-</u> difference-between-a-scientific-law-and-<u>theory-matt-anticole#watch</u> Fact vs. Theory vs. Hypothesis vs. Law... EXPLAINED! (Video)



https://youtu.be/lqk3TKuGNBA

Science at multiple levels (Webpage)



https://undsci.berkeley.edu/article/howscie nceworks_19

APPENDIX 2

Using the Kinetic Molecular Theory of Gases to explain the behaviour of gases

1. Gas Pressure

Gas molecules are in constant random motion within the container that the gas is in and collide with the walls of the container. When they do so, they exert a force which results in

pressure being set up $\left(\text{pressure} = \frac{\text{force}}{\text{area}} \right)$.

2. Boyle's Law

When the volume of the gas decreases, the distances between the gas molecules and the walls of the container are shorter. Hence, molecules collide with the walls more **frequently**, leading to greater pressure being exerted *by* the gas.

When the pressure exerted *on* the gas increases, gas molecules are forced closer together, leading to a decrease in volume.

3. Charles' Law

When temperature of the gas increases, the gas molecules have higher kinetic energy and move faster. They hit the walls of the container more frequently and with greater force, resulting in greater pressure being exerted by the gas. This results in the walls of the container moving outwards (i.e. increase in volume) until constant pressure is restored.

4. Avogadro's Law/ Dalton's Law of Partial Pressure

The pressure exerted by the gas depends on the number of collisions per second, which is in turn proportional to the number of gas molecules present.

APPENDIX 3

Trend of the *pV*–*p* graph for a real gas

As we have discussed in the previous sections, the behaviours of real gases do not always obey the ideal gas equation, pV = nRT.



In 1873, Dutch physicist Johannes van der Waals successfully accounted for the behaviour of real gases by introducing two additional terms to the ideal gas equation, resulting in the **van der Waals equation**:



The constants *a* and *b* differ for different gases: *a* correlates with the strength of the intermolecular forces between the gas particles, while *b* correlates with the size of the molecules. The smaller the constants are for a real gas, the more ideal it is. The van der Waals constants for some gases are listed in the table below:

gas	<i>a I</i> Pa m ⁶ mol ⁻²	<i>b</i> / 10 ⁻⁵ m³ mol ⁻¹
water vapour	0.5537	3.049
ammonia	0.4225	3.713
carbon dioxide	0.3658	4.286
carbon monoxide	0.1472	3.948
nitrogen	0.1370	3.87
hydrogen	0.02453	2.651
helium	0.00346	2.38

Do these values concur with what you have learnt so far about how much different gases deviate from ideality?