The Gaseous State

Content
- Ideal gas behaviour and deviations from it
- $pV = nRT$ and its use in determining a value for M_r
- Dalton's Law and its use in determining the partial pressures of gases in a mixture

Learning Outcomes required for H2 (9729) Chemistry:
Candidates should be able to
- state and use the general gas equation $pV = nRT$ in calculations, including the determination of M_r
- use Dalton's Law to determine the partial pressure of gases in a mixture
- state basic assumptions of the kinetic theory as applied to an ideal gas
- explain qualitatively in terms of intermolecular forces and molecular size:
  - the conditions necessary for a real gas to approach ideal behaviour;
  - the limitations of ideality at very high pressures and very low temperatures

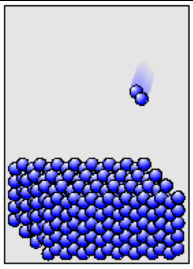
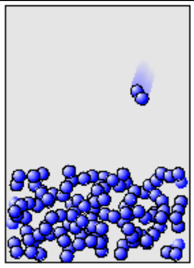
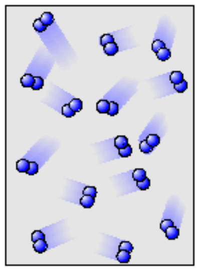
References:
- Chemistry in Context by *G Hill & J Holman*
- Chemistry for Advanced Level by *Peter Cann & Peter Hughes*

Instructions to Students

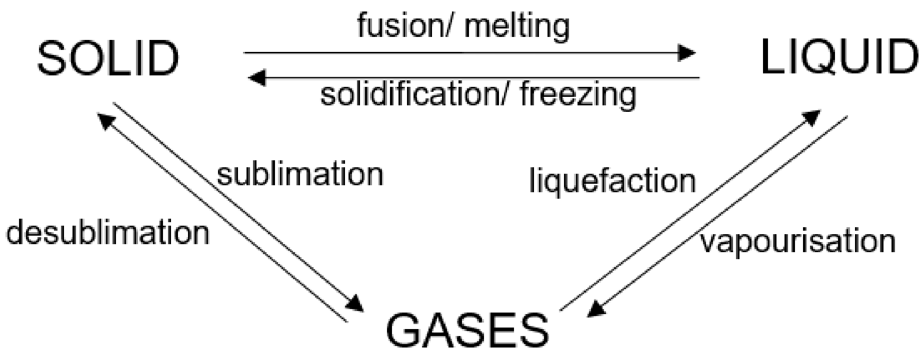
Period	Activities
13 March – 9 April	Self-Directed Learning Access: 1) Google classroom for Checkpoint suggested answers 2) SLS lesson - "Ideal and Real Gases" 3) SLS lesson - "Ideal Gas Laws"
10 April – 14 April	1 Tutorial for topic discussion
17 April – 21 April	Class Test

1 Introduction

1.1 The three states of matter

Physical State	**Solid**	**Liquid**	**Gas**
Attractive forces between particles	Strong	Intermediate	Weak
Motion of particles	Vibrate and rotate about fixed positions (no translation motion)	Move freely throughout liquid (vibrate, rotate and translate)	Move freely within the container (vibrate, rotate and translate)
Space between particles	Negligible	Small	Large
Order of Packing of Particles	 Regular packing in crystals	 Loose clusters of particles	 No order; far apart
Comparing Properties of Solids, Liquids and Gases			
Volume	Fixed	Fixed	Particles will spread evenly throughout any container, thus volume is indefinite and shape is that of the container.
Shape	Fixed	Takes the shape of the container but may not occupy it completely	
Relative Density	High	Moderate	Low
Relative Compressibility	Negligible or almost nil	Very slight	Very easily compressed

1.2 Relationship between the three states of matter



2 The Gas Laws

The properties of gases were studied as their behaviour deviated very much from that of solids and liquids. Scientists conducted experiments and derived gas laws from their results.

The gas laws were formulated by measuring how the volume of a sample of gas varies with amount of gas, pressure and temperature.

Success Criteria:
- Be able to state and use the general gas equation $pV = nRT$ in calculations, including the determination of M_r

2.1 Parameters of Gas Laws

Four parameters are needed to define the state of a gas.

Note:
SI units should be used consistently in calculations involving gas equations.

Parameter	Symbol	SI unit	Conversion
Amount of the gas	n	mol	-
Temperature of the gas	T	K (Kelvin)	x °C = (x + 273) K
Volume of the gas	V	m^3	$1 m^3 = 1 \times 10^3 dm^3$ $1 m^3 = 1 \times 10^6 cm^3$
Pressure of the gas	p	Pa (Pascal)	1 Pa = 1 Nm^{-2} 1 atm = 101325 Pa 1 bar = 10^5 Pa

2.2 Avogadro's Law

At constant temperature and pressure, the **volume** of an ideal gas is **proportional** to the amount of gas.

$$\boxed{V \propto n}$$

At standard room temperature and pressure (s.t.p):
10^5 Pa (1 bar) and 273 K (0 °C)
- 1 mol of gas occupies a volume of 22.7×10^{-3} m^3 (22.7 dm^3)
- 2 mol of gas occupies a volume of 45.4×10^{-3} m^3 (45.4 dm^3)

At room temperature and pressure (r.t.p):
101325 Pa (1 atm) and 293 K (20 °C)
- 1 mol of gas occupies a volume of 24×10^{-3} m^3 (24.0 dm^3)
- 2 mol of gas occupies a volume of 48×10^{-3} m^3 (48.0 dm^3)

2.3 Boyle's Law

At constant temperature, the **volume** of a given mass of ideal gas is **inversely proportional** to the applied **pressure**.

$$\boxed{V \propto \frac{1}{p}}$$

2.4 Charles' Law

Note:
If the temperature data is given in degree Celsius, you must convert the value into Kelvin. Please refer to the table in 2.1 for the unit conversion.

The **volume** of a given mass of ideal gas at constant pressure is **directly proportional** to its **temperature** expressed in **Kelvin**.

$$\boxed{V \propto T}$$

2.5 The Combined Gas Law

Combining Avogadro's Law ($V \propto n$), Boyle's Law ($V \propto \frac{1}{p}$) and Charles' Law ($V \propto T$), we have:

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

$$V = \text{constant} \times \frac{nT}{p}$$

$$\boldsymbol{\frac{pV}{nT} = \textbf{constant}}$$

The exact value of the constant depends on the amount of gas (n).

Thus for 1 mol of an ideal gas, the constant
(i) is denoted by the symbol ***R***
(ii) is known as the **molar gas constant**
(iii) has a value of **8.31 J K^{-1} mol^{-1}** (in SI unit)

2.6 The Ideal Gas Equation

Note:
The value of ***R*** is taken to be **8.31 J K^{-1} mol^{-1}** in the ideal gas equation, the following units must be used.

Parameter	Units
p	Pa
V	m^3
T	K

Since for 1 mol of gas, $\boldsymbol{\frac{pV}{T} = R}$

Therefore for n mol of gas, $\boldsymbol{\frac{pV}{T} = nR}$

Rearranging, we have the ideal gas equation

$$\boxed{\boldsymbol{pV = nRT}}$$

Worked Example 1

At 150×10^3 Pa, a fixed mass of an ideal gas occupies 300×10^{-6} m^3 at temperature T. What will be its volume at 250×10^3 Pa at the same temperature?

Using the ideal gas equation, $p_1V_1 = nRT$ – (1) and $p_2V_2 = nRT$ – (2)
Given a fixed amount of ideal gas at constant temperature T, (1) = (2)

$$p_1V_1 = p_2V_2$$
$$(150 \times 10^3)(300 \times 10^{-6}) = (250 \times 10^3)(V_2)$$
$$\mathbf{V_2 = \underline{1.80 \times 10^{-4}\ m^3}}$$

Worked Example 2

A sample of neon gas occupies a volume of 4.0×10^{-3} m^3 at 50 °C. Determine the new volume occupied by the neon gas when the sample is heated to 80 °C at constant pressure.

Note:
You must always convert the given temperature in degree Celsius to Kelvin scale.

Using the ideal gas equation, $pV_1 = nRT_1$ – (1) and $pV_2 = nRT_2$ – (2)
Given a fixed amount of Neon (g) at constant pressure, $\frac{(1)}{(2)}$

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}$$
$$\frac{4.0 \times 10^{-3}}{V_2} = \frac{50 + 273}{80 + 273}$$

$$\mathbf{V_2 = \underline{4.37 \times 10^{-3}\ m^3}}$$

Worked Example 3

Note:
We can also use the volume in dm^3 for calculations

$$\frac{p_1V_1}{p_2V_2} = \frac{T_1}{T_2}$$

$$\frac{(1.0 \times 10^5)(6\ dm^3)}{(0.45 \times 10^5)(V_2)} = \frac{(22 + 273)}{(-21 + 273)}$$

$$\mathbf{V_2 = \underline{11.4\ dm^3}}$$

An inflated balloon has a volume of 6 dm^3 at sea level (1.0×10^5 Pa) and is allowed to ascend in altitude until the pressure is 0.45×10^5 Pa. During ascent, the temperature falls from 22 °C to –21 °C. Calculate the volume of the balloon at its final altitude.

Using the ideal gas equation, $p_1V_1 = nRT_1$ – (1) and $p_2V_2 = nRT_2$ – (2)
Given a fixed amount of gas inside the balloon, $\frac{(1)}{(2)}$

$$\frac{p_1V_1}{p_2V_2} = \frac{T_1}{T_2}$$
$$\frac{(1.0 \times 10^5)(6 \times 10^{-3})}{(0.45 \times 10^5)(V_2)} = \frac{(22 + 273)}{(-21 + 273)}$$

$$\mathbf{V_2 = \underline{11.4 \times 10^{-3}\ m^3}}$$

Worked Example 4

Tennis balls are usually filled with air or nitrogen gas to a pressure above atmospheric pressure to increase their “bounce”. If a particular tennis ball has a volume of 144 cm^3 and contains 0.33 g of nitrogen gas, what is the pressure inside the ball at 24 °C?

Amount of N_2 gas (n) = $\frac{0.33}{28.0}$ = 1.179×10^{-2} mol
Using the ideal gas equation, $pV = nRT$
$p(144 \times 10^{-6}) = (1.179 \times 10^{-2})(8.31)(24 + 273)$

***p* = <u>202 × 10³ Pa</u>**

Learning points:
For calculations involving

1) **<u>only one</u>** $pV = nRT$ equation, **<u>ALL</u>** parameters need to be changed to SI units when doing calculation.

2) **<u>two</u>** $pV = nRT$ equations, **<u>only temperature</u>** parameters need to be changed to SI units when doing calculation. Unit conversions for pressures (atm or kPa to Pa) and volumes (cm^3 or dm^3 to m^3) would be cancelled out when combining **<u>two</u>** $pV = nRT$ equations.

<u>Checkpoint 1</u> 

1) The compressed–air cylinder of a particular scuba diving outfit has an internal volume of 1.88 dm^3, and contains air at a pressure of 1.75×10^5 Pa. Calculate the volume of air that is available at a pressure of 1.01×10^5 Pa at the same temperature. State your assumption.

2) A syringe containing 30.2 cm^3 of an ideal gas at 20 °C, with a pressure of 1 atm acting on the plunger, was cooled to −50 °C. To what value will the volume decrease if the pressure remains constant?

Checkpoint 1 (Continued)

3a) A meteorological balloon is to be filled with helium at atmospheric pressure. What will be the volume of the balloon if it is to hold all the helium gas from a 25.0 dm^3 gas cylinder at 150 atm? Assume constant temperature.

3b) The balloon will burst if its volume exceeds 4000 dm^3. If the balloon in (a) was filled at 15 °C, what is the maximum temperature that the balloon can withstand without bursting at atmospheric pressure?

2.7 Further manipulations of the Ideal Gas Equation

The Ideal Gas Equation can be manipulated to determine properties such as concentration, molar mass (or relative molecular mass) and density of a gas.

To determine the **concentration** of a gas, the Ideal Gas Equation is rearranged as follows:

$$pV = nRT$$

$$p = \frac{n}{V}RT$$

To determine the **density,** ρ**,** of a gas, the Ideal Gas Equation can be rearranged as follows:

$$pV = nRT$$

$$pV = \frac{m}{M}RT$$

(where m is mass in g, and M is the molar mass in g mol^{-1})

$$p = \frac{m}{V} \times \frac{RT}{M}$$

$$p = \rho \times \frac{RT}{M}$$

Worked Example 5

A 0.258 g gas sample at 182 °C and 100 kPa has a volume of 115 cm^3. Determine the relative molecular mass of the gas.

$$pV = nRT$$

$$pV = \frac{m}{M}RT$$

$$(100 \times 10^3)(115 \times 10^{-6}) = \left(\frac{0.258}{M}\right)(8.31)(182 + 273)$$

$$M = 84.8 \text{ g mol}^{-1}$$

Thus, the relative molecular mass, M_r = 84.8 (no units)

Checkpoint 2

1) **(i)** Calculate the volume occupied by 2.00 g of carbon monoxide at 20 °C under a pressure of 6250 Pa.

(ii) What will its volume be if its temperature is raised to 40 °C, keeping the pressure constant?

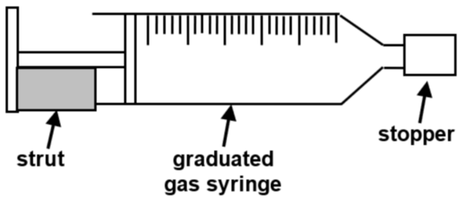
2) What mass of zinc is required to produce 2.00 dm^3 of hydrogen at 15 °C and 100 kPa by reacting with excess acid?

3) A gas is found to have a density of 2.65 g dm^{-3} at 24 °C and 102 kPa. Determine the relative molecular mass of the gas.

Note: It is necessary to carry out the unit conversion to SI units for density before calculation.

2.8 Experimental setups for determining molar masses of gases and volatile liquids

1) Gases

Direct Weighing Method	**Brief Procedure:**
	1. Weigh a known volume of the gas sample at a given temperature and pressure. 2. Determine the molar mass, M, using the Ideal Gas Equation: $\boldsymbol{pV = nRT}$ where $\boldsymbol{n = \frac{m}{M}}$

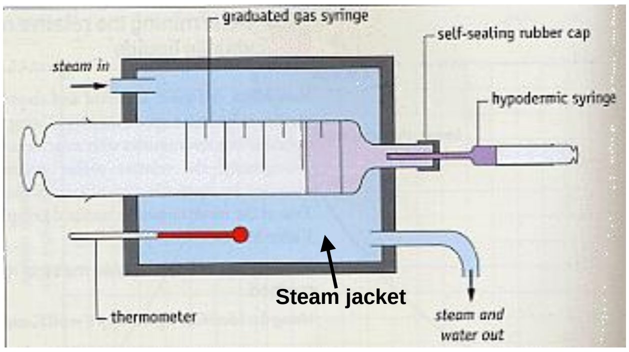
Data collected

Volume of gas	/cm^3	*x*
Mass of 'empty' syringe	/g	
Mass of syringe with *x* cm^3 of gas	/g	
Atmospheric pressure	/kPa	
Temperature of room	/°C	

2) Volatile Liquids

Use steam to vapourise liquid

(This method is not suitable for liquids with boiling point above 80 °C.)



Brief Procedure:

1. Weigh a known volume of the gas sample at a given temperature and pressure.
2. Determine the molar mass, M, using the Ideal Gas Equation:

$\boldsymbol{pV = nRT}$ where $\boldsymbol{n = \frac{m}{M}}$

Data collected

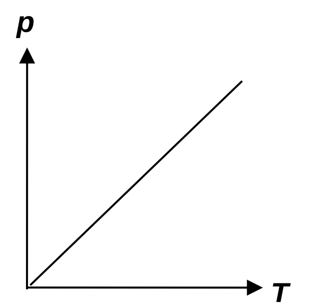
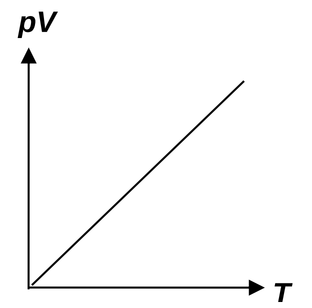
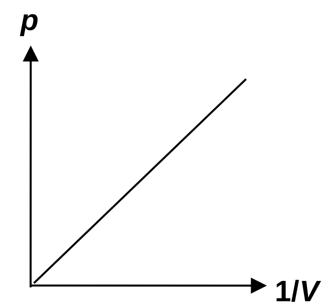
Mass of hypodermic syringe and liquid before injection into gas syringe	/g	
Mass of hypodermic syringe and liquid after injection into gas syringe	/g	
Final reading on gas syringe	/cm^3	
Initial reading on gas syringe	/cm^3	
Volume of gas	/cm^3	
Atmospheric pressure	/kPa	
Temperature of steam jacket	/°C	

2.9 Graphical Representations of the Relationship Between Parameters in the Ideal Gas Equation.

Generally there are only 3 possible relationships between the parameters:

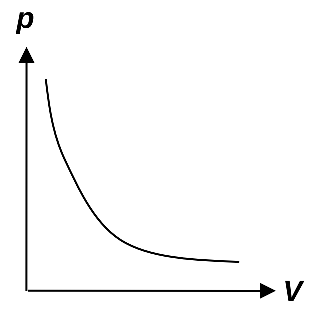
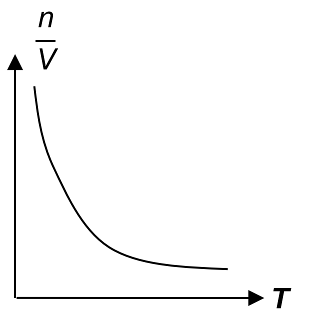
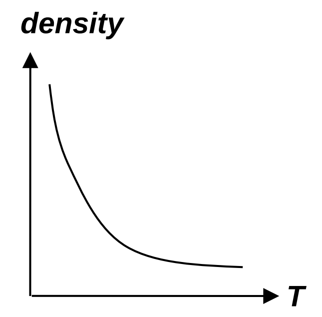
1) Directly proportional, $Y = mX$

Boyles Law

		
$\boxed{p} = \frac{nR}{V}\boxed{T}$	$\boxed{pV} = nR\boxed{T}$	$\boxed{p} = nRT\boxed{\frac{1}{V}}$
n and *V* constant	*n* constant	*n* and *T* constant

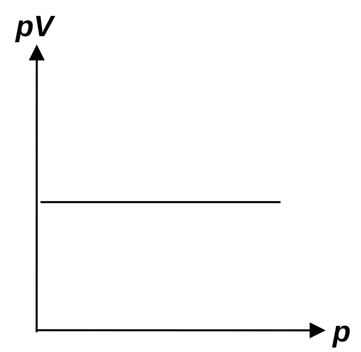
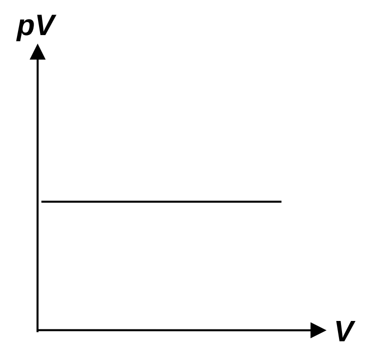
2) Inversely proportional, $Y = \frac{m}{X}$

Charles Law

		
$\boxed{p} = \frac{nRT}{\boxed{V}}$	$\boxed{\frac{n}{V}} = \frac{p}{R\boxed{T}}$	$\boxed{\rho} = \frac{pM}{R\boxed{T}}$
n and *T* constant	*p* constant	*p* constant

3) The parameters (*Y* & *X*) are independent of each other, $Y = c$

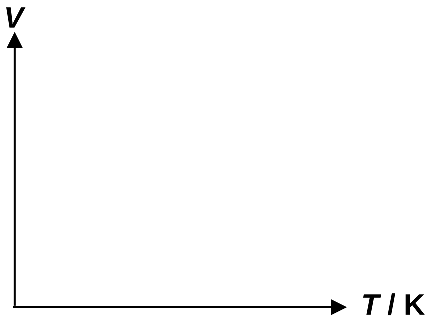
Note:
The value of *pV* is only affected by *n* and *T*, hence even when *p* OR *V* is varied, the value of *pV* still remains constant.

	
$\boxed{pV} = nRT$	$\boxed{pV} = nRT$
n and *T* constant	*n* and *T* constant

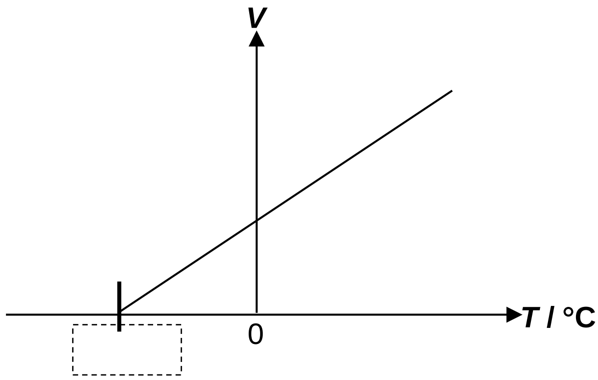
Checkpoint 3 

1) Illustrate the relationship between volume and temperature of a fixed mass of gas at constant pressure on the graph (i) and give a value for the *x*-intercept in graph (ii).

(i)

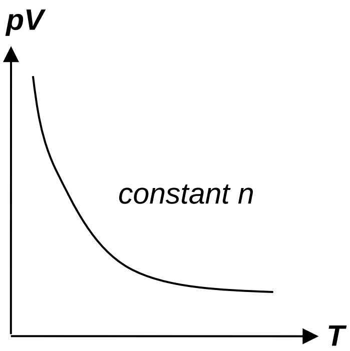


(ii)

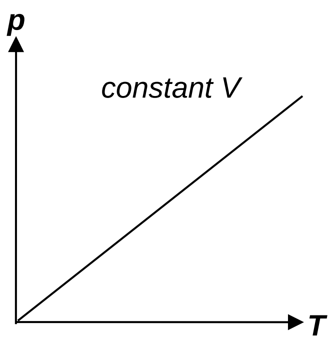


2) Which graph(s) represent the behaviour of a fixed amount of an ideal gas?

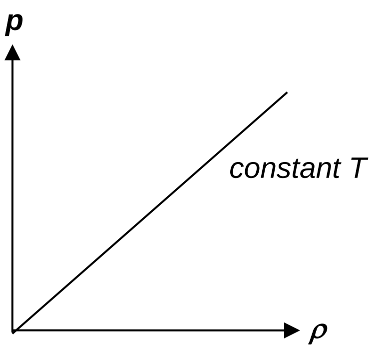
(A) Graph of pV vs T



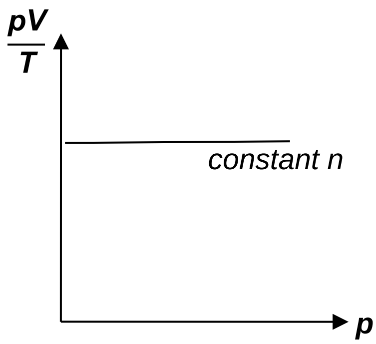
(B) Graph of p vs T



(C) Graph of p vs ρ



(D) Graph of $\frac{pV}{T}$ vs p

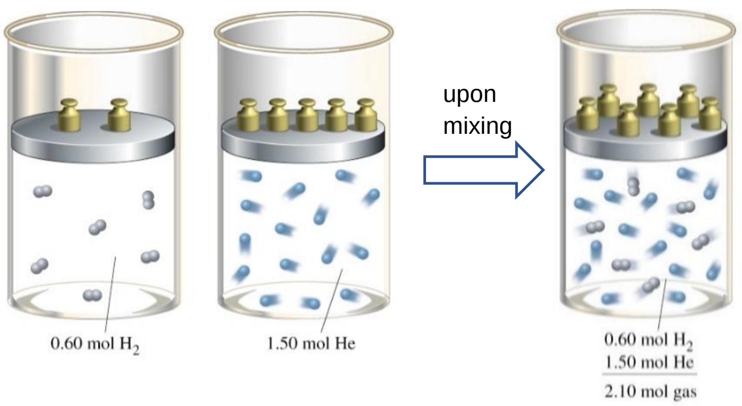


Answer :

3 Mixture of Gases

When gases, which **do not react chemically with each other**, are mixed in a vessel, each gas will behave independently of the other gases. The contribution which each gas makes to the total pressure in the vessel is known as the **partial pressure**.

Mixing of hydrogen and helium gas under same volume



The **partial pressure** of a gas in a mixture is the pressure the gas would exert if it alone occupies the same volume of the container at the same temperature. Therefore, partial pressure of gas $\propto$ amount of gas in a certain volume of vessel and temperature. There is only partial pressure, no partial volume.

Success Criteria:

- To be able to use Dalton's Law to determine the partial pressure of gases in a mixture

3.1 Dalton's Law of Partial Pressure

At constant temperature, the **total pressure** of a mixture of gases which do not react chemically with each other is equal to the **sum of the partial pressures** of the gases present in the mixture.

$$p_{total} = p_A + p_B + p_C + p_D + \ldots$$

where p_{total} = total pressure of the gaseous mixture,

p_A, p_B, p_C, p_D = partial pressures of gases ***A***, ***B***, ***C***, ***D*** respectively

3.2 Relationship between Partial Pressure and Mole Fraction

In a mixture of three gases, ***A***, ***B*** and ***C***, with partial pressures p_A, p_B, p_C, the total pressure can be calculated from the Ideal Gas Equation.

$$p_{total}V = n_{total}RT \quad \text{----------(1)}$$

Similarly, the partial pressure of each gas in the mixture, e.g. gas A, can also be calculated using the ideal gas equation.

$$p_AV = n_ART \quad \text{----------(2)}$$

By rearranging equations 1 and 2, $\frac{RT}{V}$ can be expressed in terms of *P* and *n* and $\frac{RT}{V}$ is a constant value.

From equation 1: $\frac{p_{total}}{n_{total}} = \frac{RT}{V}$

From equation 2: $\frac{p_A}{n_A} = \frac{RT}{V}$

Hence, $\frac{p_A}{n_A} = \frac{p_{total}}{n_{total}} = \frac{RT}{V}$

$$\frac{p_AV}{p_{total}V} = \frac{n_ART}{n_{total}RT}$$

$$\boxed{p_A = \left(\frac{n_A}{n_{total}}\right) \times p_{total}}$$

where $\left(\frac{n_A}{n_{total}}\right)$ is the mole fraction of gas ***A*** **i**n the mixture

The **mole fraction** of a component (i.e gas ***A)*** in a mixture is defined as the **ratio** of the number of moles of gas ***A*** to the total number of moles of gas present (i.e gases ***A***, ***B*** and ***C***) in the mixture.

$$\boxed{\textbf{Mole fraction of A} = \frac{n_A}{n_{total}} = \frac{n_A}{n_A + n_B + n_C}}$$

Worked Example 6

A solid hydrocarbon is burned in air in a closed container, producing a mixture of gases having a total pressure of 3.34 atm at 298 K. Analysis of the mixture at 298 K shows it contains 0.340 g of water vapour, 0.792 g of carbon dioxide, 0.288 g of oxygen, 3.790 g of nitrogen and no other gases.

Calculate the mole fraction and partial pressure of carbon dioxide in this mixture.

$$n_{total} = n_{H_2O} + n_{CO_2} + n_{O_2} + n_{N_2} = \frac{0.340}{18.0} + \frac{0.792}{44.0} + \frac{0.288}{32.0} + \frac{3.790}{28.0}$$

$$= 0.1812 \text{ mol}$$

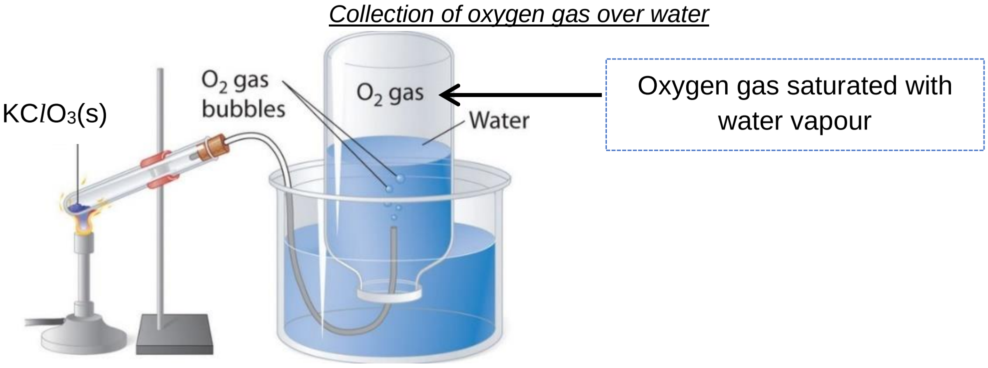
$$\text{Mole fraction of } CO_2 = \frac{n_{CO_2}}{n_{total}} = \frac{0.018}{0.1812} = \underline{0.0993}$$

$$p_A = \frac{n_{CO_2}}{n_{total}} \times p_{total} = 0.0993 \times 3.34 = \underline{0.332 \text{ atm}}$$

3.3 Applications of Dalton's Law of Partial Pressures

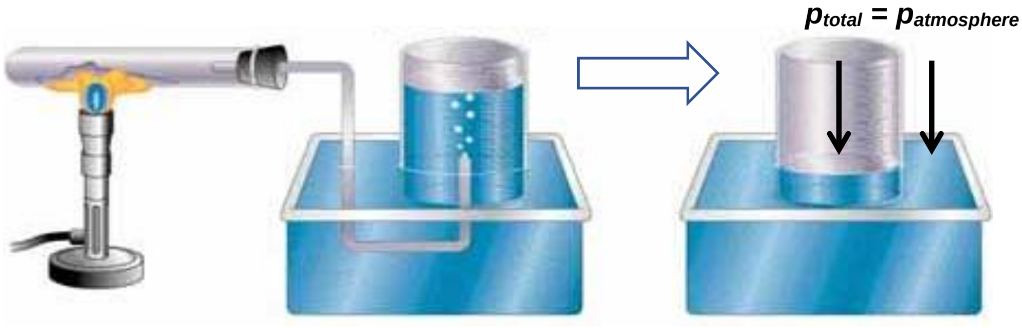
When a gas is collected over water, there is water vapour mixed in with the gas. The gas is said to be saturated with water vapour.

Collection of oxygen gas over water



The water level inside the gas measuring tube is then adjusted to be equal to the external water level. Thus, the total pressure of the gases inside the measuring tube is equal to the atmospheric pressure (i.e external pressure).

Gas volume measurement



Dalton's Law can then be used to determine the partial pressure of the dry gas.

$$p_{total} = p_{dry\ gas} + p_{water\ vapour} = p_{atmosphere}$$

$$p_{dry\ gas} = p_{atmosphere} - p_{water\ vapour}$$

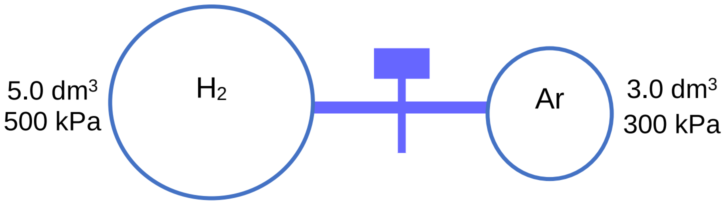
Checkpoint 4

1) A 10.0 dm^3 container contains 2.00 g of hydrogen and 7.00 g nitrogen. If the temperature of the mixture is 30 °C, calculate

(i) the pressure of the mixture

(ii) the partial pressure of each gas

2) A 5.0 dm^3 cylinder containing hydrogen at a pressure of 500 kPa is connected to a 3.0 dm^3 cylinder containing argon at a pressure of 300 kPa and sufficient time is allowed for the gases to mix completely. Calculate the partial pressure of the two gases and the total pressure in the joined cylinders. The system is maintained at a constant temperature. State any assumptions you made in your calculation.



Checkpoint 4 (Continued)

3) A sample of $KClO_3$ is partially decomposed, producing oxygen gas that is collected over water. The volume of gas collected is 0.250 dm^3 at 25 °C and 1.02×10^5 Nm^{-2} total pressure.

$$2KClO_3\ (s) \rightarrow 2KCl\ (s) + 3O_2\ (g)$$

The vapour pressure of water at various temperatures is presented in the following table.

Temperature/ °C	Vapour Pressure/ $\times 10^3$ Nm^{-2}
15	1.70
20	2.33
25	3.16
30	4.23

(i) What is the amount of oxygen collected?

(ii) What is the mass of $KClO_3$ decomposed?

4 The Kinetic Theory of Gases

The Ideal Gas Equation is an empirical law based on experimental observations and describes the behaviour of a gas in macroscopic terms. The Kinetic Theory of Gases is a simple model that attempts to explain the behaviour of the particles in the gases.

Success Criteria:
- To be able to state basic assumptions of the kinetic theory as applied to an ideal gas

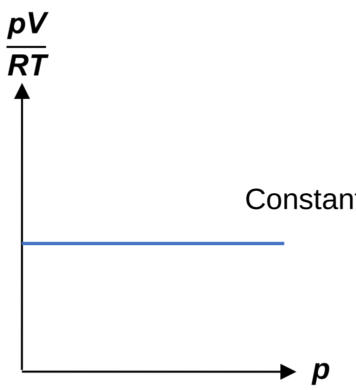
4.1 Assumptions of the Kinetic Theory

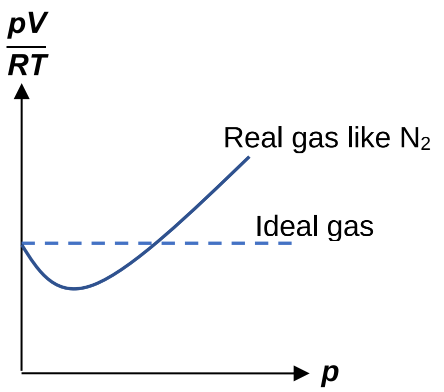
Note:
Point 1 and 2 are the **main** assumptions made by the kinetic theory about Ideal Gases.

1. A gas is composed of tiny particles that have **negligible volumes(size)** compared to the volume of the container.
2. There are **negligible intermolecular forces** between the gas particles.
3. Collisions between gas particles are perfectly elastic (no net loss in kinetic energy).

4.2 The behaviour of real gases

A gas is said to exhibit ideal gas behaviour if it obeys the ideal gas equation exactly under any conditions. A plot of $\frac{pV}{RT}$ versus p will give a horizontal straight line.



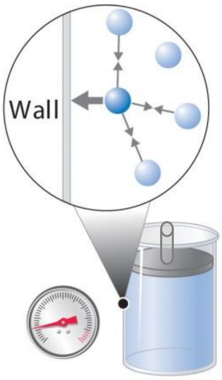


However, real gases do not behave ideally under all conditions. When experiments are performed on a real gas such as nitrogen, the plot of $\frac{pV}{RT}$ versus p does not yield a straight line. Instead, the graph shows a deviation from the ideal plot.
(The deviation from the ideal plot will be covered in section 4.2.3.)

The reasons for real gases showing deviation from ideal behaviour are:
1. There are **intermolecular forces** between the gas particles.
   (refer to section 4.2.1)

2. The gas particles take up space and do occupy finite **volumes**.
   (refer to section 4.2.2)

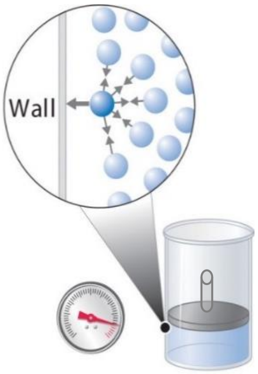
4.2.1 The effect of intermolecular forces of attraction between gas particles on the measured gas pressure



In this container (with a movable piston), the gas particles are too far apart to interact.

The pressure of the gas is the force exerted on the wall of the container when the gas particles collide with the wall.

It is equal to the external pressure that is exerted to hold the piston in place.

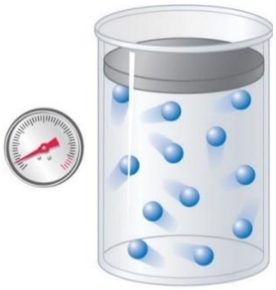


As $p \propto \frac{1}{V}$, when the volume of the container is halved (by pushing down the piston), the pressure of the gas is expected to double.

However, the gas particles are now close enough to interact. The intermolecular forces of attraction between gas particles will lower the force of collision of particles with the wall of container.

So the final pressure is less than double of the original pressure.

4.2.2 The effect of molecular volume on the measured gas volume



In this container (with a movable piston), the gas particles are far apart.

The free volume of space in which the gas particles can move around is equal to the volume of the container as volume taken up by the gas molecules is negligible compared to the volume of the container.

When the external pressure exerted on the container is doubled, the volume of container is halved.

However, the gas particles are very close together. The free volume of space in which the gas particles can move around is now less than the volume of the container as volume taken up by the gas molecules becomes non-negligible compared to the volume of the container.

Success Criteria:

- explain qualitatively in terms of intermolecular forces and molecular size:
  - the conditions necessary for a real gas to approach ideal behaviour;
  - the limitations of ideality at very high pressures and very low temperatures

4.2.3 Factors affecting deviation of real gas from ideal gas behaviour

A real gas would **behave most like an ideal gas** (or approach ideal behaviour) **at low pressure and high temperature**.

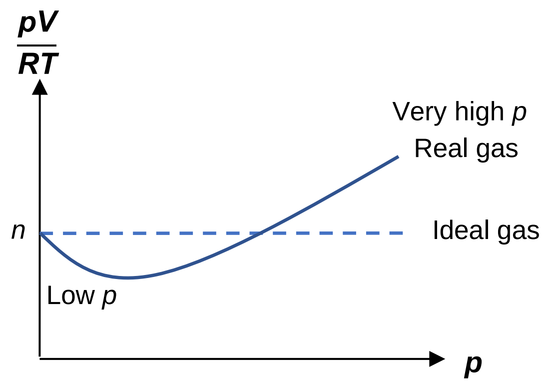
It deviates greatly from ideal behaviour at high pressure and low temperature.

The extent and degree of deviation of real gas from ideal gas behaviour depends on:

a) pressure
b) temperature
c) the nature of the gas (type of intermolecular forces and molecular size to compare the volume of a gas molecule)

a) Pressure

Graph of $\frac{pV}{RT}$ vs p of a real gas vs ideal gas



At low pressure,

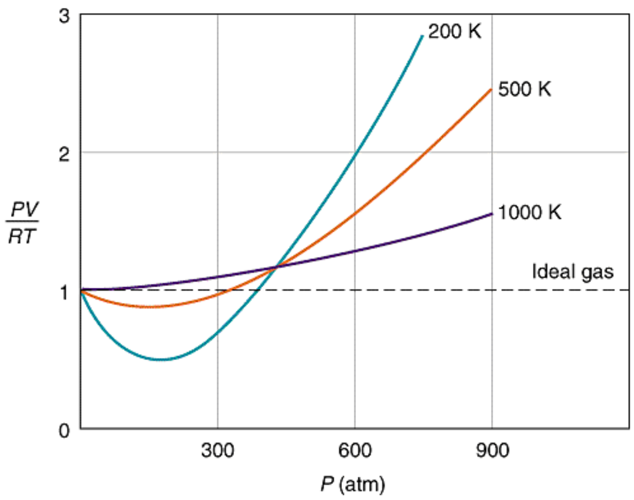
- The gas particles are far apart so the intermolecular forces of attraction between them are negligible.
- The total volume of the gas particles is negligible compared to the volume of the container.

Deviations at high pressure

- The gas particles are very close. The total volume of the gas particles is no longer negligible compared to the volume of the container.

b) Temperature

Graph of $\frac{pV}{RT}$ vs p of 1 mol of a real gas at varied temperature



<u>At high temperature</u>

- The gas particles are moving rapidly and they have sufficient kinetic energy to overcome the intermolecular forces of attractions between them. So the intermolecular forces of attraction can be considered as negligible.

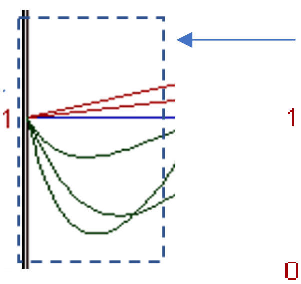
<u>Deviation at low temperature</u>

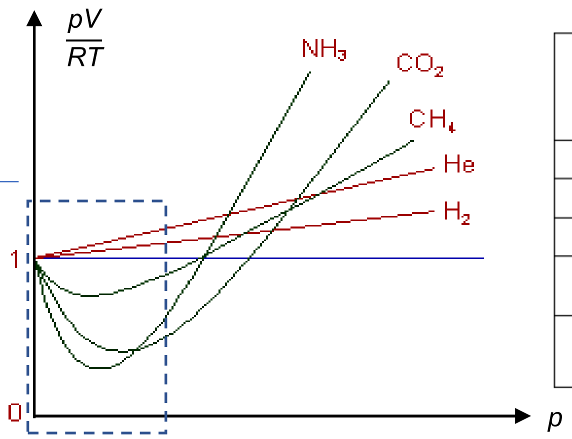
- The gas particles have lower kinetic energy and are insufficient to overcome the intermolecular forces of attractions. The intermolecular forces of attractions between them become significant and are no longer negligible.

c) Nature of the Gas

Graph of $\frac{pV}{RT}$ vs p of under constant temperature for 1 mol of different gases

Note:
Focus on the low pressure region to compare the degree of deviation from ideal gas (horizontal line)





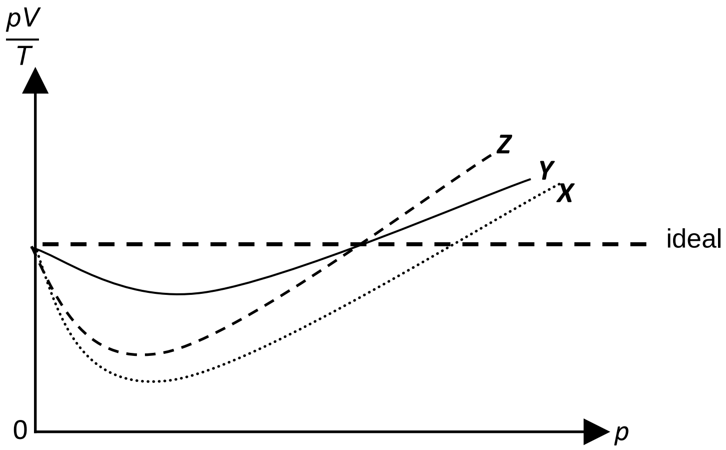
Gas	Nature of intermolecular forces	Strength of intermolecular forces
H_2	Instantaneous dipole – induced dipole (id-id)	increases
He		
CH_4		
CO_2		
NH_3	Hydrogen bonding	

- The stronger the intermolecular forces of attraction of a gas, the greater the deviation from ideal gas behaviour.

- Gases with larger molecular mass indicate a larger molecular size, therefore stronger instantaneous dipole-induced dipole interactions between molecules.

Checkpoint 5 

1) The following graphs are drawn for equal amounts of an ideal gas and three real gases at constant temperature.
Answer the guiding questions and suggest the identities of real gases **X**, **Y** and **Z**.



	X	Y	Z
A	SO_3	CH_4	N_2
B	N_2	SO_3	CH_4
C	CH_4	N_2	SO_3
D	CH_4	SO_3	N_2

Guiding questions:

a) Use the graphs to determine the extent of deviation from ideal gas behaviour for X, Y and Z at low pressures.

b) Using chemical bonding knowledge, identify the type of intermolecular forces of attraction for each gas in the options.

c) Compare the strength of intermolecular forces of attraction for the three gases.

d) Apply “the greater the intermolecular forces of attraction, the greater the deviation from ideal gas behaviour” and select the option showing the correct identities.

For Interest:

The van der Waals Equation

Engineers and scientists who work with gases at high pressures often cannot use the Ideal Gas Equation to predict the pressure – volume properties of gases. This is because the behaviour of the gases deviating from ideal behaviour is too large.

$$\left(p + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

(where p is the measured pressure and V is the volume of container)

Correction for forces of attraction between molecules

Correction for volume of molecules

The van der Waals Equation takes in to account the attractive forces present in real gas and the finite volume occupied by the gas molecules. This equation is able to predict the behaviour of real gases better.

Real – life Applications of Gas Laws

(A) Airbags in Modern Cars

Most modern cars are equipped with an airbag, which inflates in a few microseconds during sudden impacts. This protects the driver and the passenger seated in the front seat, who may otherwise be thrown against the steering wheel or dashboard.

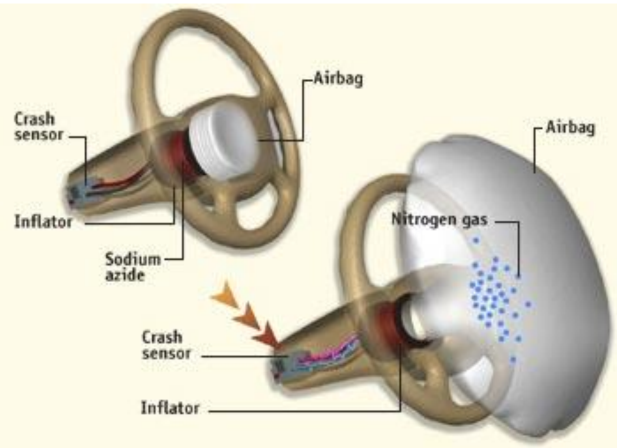
So what enables the deflated airbag to inflate within microseconds during sudden impacts?

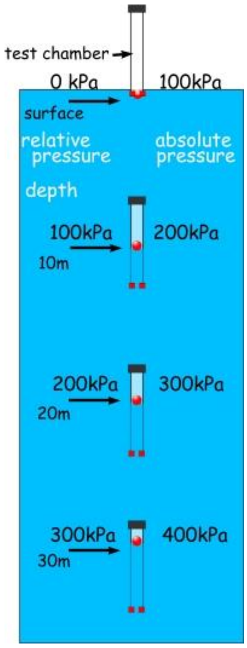
There is a solid pellet of sodium azide, NaN_3, in the airbag.
During a sudden impact, NaN_3 will be ignited and undergo decomposition, generating a large amount of nitrogen gas, N_2.

$$2NaN_3\,(s) \rightarrow 2Na\,(s) + 3N_2\,(g)$$

The N_2 gas produced fills the entire airbag, inflating it in the process.

At constant temperature and pressure, a mass of 124 g of NaN_3 is needed to fill a 70 dm^3 airbag with N_2 gas. Thus it is easier to carry the NaN_3 pellet around than a bag full of air.





(B) Scuba diving

A recreational dive in Maldives under the supervision of a dive professional allows a diver to explore the underwater up to a depth of 12 m.

Deep underwater, the weight of the sea water exerts much greater pressure on the body of a scuba diver than it is at the surface level. The pressure exerted by the sea water on the diver increases with depth.

Pressure exerted on a diver at the surface level	1.0 atm (p_{atm})
Pressure exerted by 10 m of sea water on a diver	1.0 atm

So a diver diving at the depth of 10 m will experience a total pressure of 2.0 atm.

As the diver begin to ascend to the surface, the volume of air in the lungs will increase with decreasing pressure, according to Boyle's Law.

So the volume of air in the lungs will double! This sudden increase in volume can rupture the membranes of the lungs. Therefore it is important that the diver ascend slowly, stopping at certain points to adjust to the falling pressure.

	Success Criteria	Relevant Tutorial Qns	What do you still struggle with? Write your queries here.
	I am able to:		
(a)	use the general gas equation $pV = nRT$ in calculations with the conversions of quantities into **SI units** (i) atm or kPa to **Pa** (ii) cm^3 or dm^3 to **m^3** (iii) °C to **K**	**DQ1, 2, 4, 5, 7, 14, 15**	
(b)	manipulate the general gas equation of $pV = nRT$, including the determination of M_r and density.	**DQ9, 13**	
(c)	sketch the graphical representation of the parameters in the ideal gas equation (i) Directly proportional, $Y = mX$ (ii) Inversely proportional, $Y = \frac{m}{X}$ (iii) Two parameters (Y and X) are independent of each other, $Y = c$	**DQ3, 11, 15**	
(d)	use Dalton's Law and mole fraction to determine the partial pressure of gases in a mixture $p_{total} = p_A + p_B$ *where* $p_A = (\frac{n_A}{n_{total}}) \times p_{total}$ $p_B = (\frac{n_B}{n_{total}}) \times p_{total}$	**DQ8, 12, 16**	
(e)	state basic assumptions of the kinetic theory as applied to an ideal gas	**Q13**	
(f)	explain qualitatively in terms of intermolecular forces and molecular size: (i) the conditions necessary for a real gas to approach ideal behaviour; (low pressure and high temperature) (ii) the limitations of ideality at very high pressures and very low temperatures	**DQ6, 10, 14**	
(g)	analyse graphical data and explain the deviations from ideal gas with reference to relative strength of intermolecular forces.		

The Gaseous State Discussion Questions

1 Which option has the largest volume at 25 °C and 101 kPa?

A 9 g of CH_4 **B** 12 g of H_2O

C 15 g of C_2H_6 **D** 20 g of CO_2

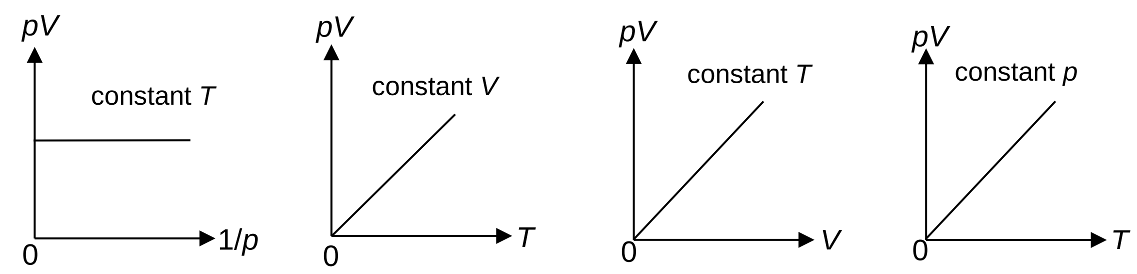
2 2.5 dm^3 of methane, CH_4, at a pressure of 1.50 atm and 4.5 dm^3 of ethane, C_2H_6, at a pressure of 1.00 atm were introduced into a 3.0 dm^3 vessel at constant temperature.

What is the final pressure of the gas mixture?

A 0.825 atm **B** 2.50 atm **C** 2.75 atm **D** 3.57 atm

3 Which graph is not a correct description of the behaviour of a fixed mass of an ideal gas?
(Note: *T* is measured in Kelvin scale.)

A **B** **C** **D**



4 A vessel contained a gas mixture of nitrogen and oxygen at 30 °C and a pressure of 200 kPa. After all the oxygen has reacted with zinc, the total pressure in the vessel was 150 kPa at 30 °C.

$$2\ Zn\ (s) + O_2\ (g) \rightarrow 2\ ZnO\ (s)$$

What was the molar ratio of nitrogen to oxygen in the mixture initially?

A 1 : 3 **B** 2 : 3 **C** 3 : 1 **D** 3 : 2

5 A container contains 1 dm^3 of gas **X** and 1 dm^3 of gas **Y** at s.t.p. Both **X** and **Y** behave as ideal gases and the relative molecular mass of **X** is 4 times that of **Y**.

Which statement is true?

A The amount of gas **X** is 4 times that of **Y**.

B The partial pressure of each gas is 50 kPa.

C The total pressure of the system will decrease when 1 mole of Ne is pumped into the container.

D The volume of the container doubles when the temperature of the system increases from 20 °C to 40 °C under constant pressure.

6 When magnesium powder reacts with dilute hydrochloric acid, hydrogen gas is evolved.

How would the volume of the hydrogen gas be affected by temperature and pressure?

A It increases with increase in temperature and is independent of pressure.

B It decreases with increase in temperature and is independent of pressure.

C It increases with increase in both temperature and pressure.

D It increases with increase in temperature and decreases with increase in pressure.

7 Which expression gives the pressure exerted by 1.6×10^{-3} mol of N_2 in a container of volume 3.0 dm^3 at 273 °C?

A $\dfrac{1.6 \times 10^{-3} \times 8.31 \times 273}{3.0 \times 10^{-6}}$ Pa

B $\dfrac{1.6 \times 10^{-3} \times 8.31 \times (273 + 273)}{3.0 \times 10^{-6}}$ Pa

C $\dfrac{1.6 \times 10^{-3} \times 8.31 \times 273}{3.0 \times 10^{-3}}$ Pa

D $\dfrac{1.6 \times 10^{-3} \times 8.31 \times (273 + 273)}{3.0 \times 10^{-3}}$ Pa

8 Flask **A** contains 1 dm^3 of helium at 2 kPa pressure and flask **B** contains 2 dm^3 of neon at 1 kPa pressure.

If the flasks are connected at constant temperature, what is the final pressure (in kPa)?

A $1\frac{1}{3}$ **B** $1\frac{1}{2}$ **C** $1\frac{2}{3}$ **D** 2

9 Which set of changes will have the greatest effect on the density of a fixed mass of an ideal gas?

	Pressure	Temperature / K
A	double	halve
B	double	double
C	halve	halve
D	halve	constant

10 Which statements are correct?

1 The density of an ideal gas at constant pressure is inversely proportional to the temperature.

2 One of the assumptions of the kinetic theory of gases is that gaseous particles are separated by distances which are large compared with their dimensions.

3 Real gases behave most nearly as an ideal gas at high pressures and low temperatures.

A 1, 2 and 3 are correct

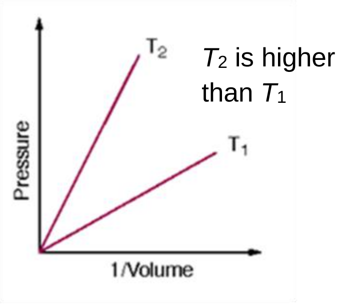
B 1 and 2 only are correct

C 2 and 3 only are correct

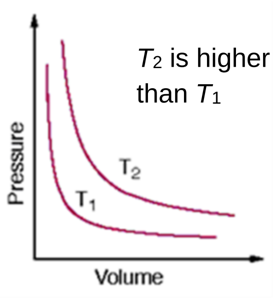
D 1 only is correct

11 Which graphs show the ideal behavior of a gas?

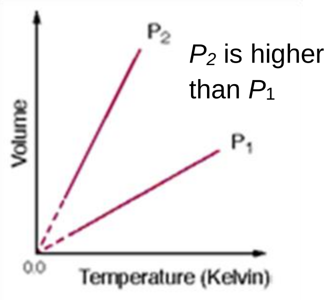
1



2



3



A 1, 2 and 3 are correct

B 1 and 2 only are correct

C 2 and 3 only are correct

D 1 only is correct

12 Two moles of oxygen and one mole of argon are contained in a cylinder with a volume of 10.0 dm^3 at 298 K.

Calculate the total pressure and partial pressure of oxygen.

13 A gas canister, used in camping stoves, contains mainly butane under sufficient pressure to cause it to liquefy partially.

(a) State two assumptions of the kinetic theory as applied to ideal gases and use these to explain whether you might expect the gas in the canister to behave as an ideal gas.

(b) The canister was connected to a gas syringe and the valve opened slightly to allow some of the gas into the syringe. It was found that 0.200 g of the gas took up a volume of 96.0 cm^3 at a temperature of 20.0 °C and a pressure of 1.01×10^5 Pa.

Calculate the average relative molecular mass of the gas mixture.

(c) Based on safety considerations, explain why the gas canister with residual butane should not be disposed of by burning.

14 Cylinders of pressurized carbon dioxide are used to produce carbonated drinks. One such cylinder has an internal volume of 2.5 dm^3, and contains 2.3 kg of carbon dioxide.

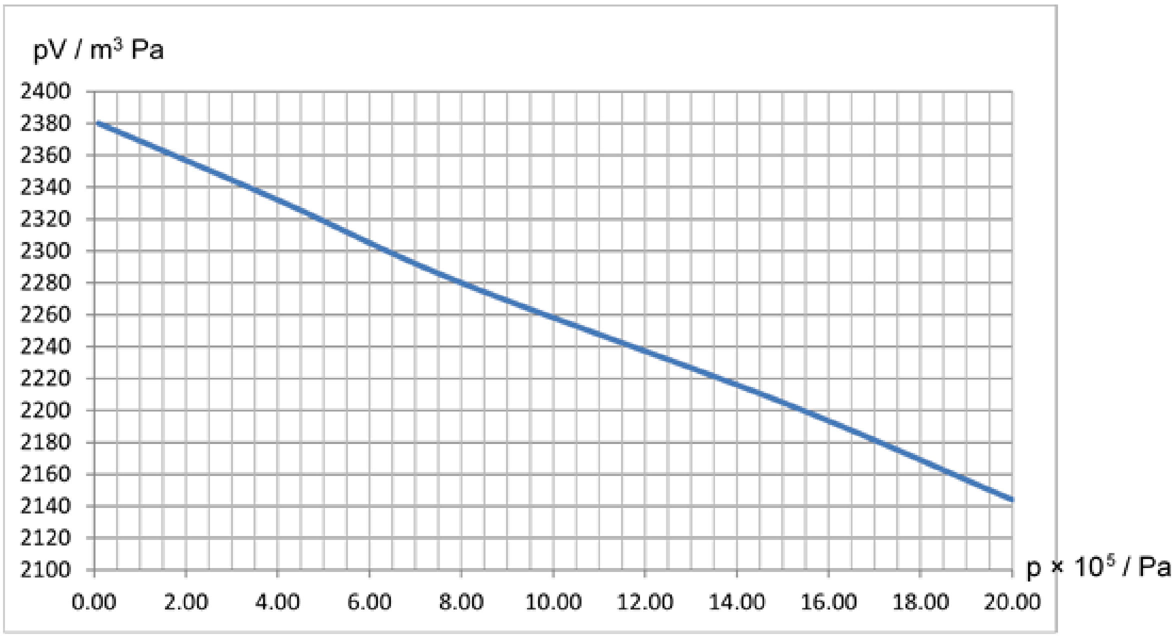
(a) Calculate the pressure (in Pa) the carbon dioxide gas would exert inside the cylinder at 25 °C.

(b) In fact, the pressure inside the cylinder is 2.2×10^7 Pa under these conditions. Explain why this differs from the pressure you calculated in (a).

(c) A 500 cm^3 can of cola has 2.0 g of carbon dioxide dissolved in it under pressure. When the can is opened, carbon dioxide is released to the atmosphere immediately till it goes flat. A saturated solution of carbon dioxide is then left in the can with a concentration of 1.5 g dm^{-3} at a pressure of 1 atm and 25 °C.

Calculate the volume of carbon dioxide that is released to the atmosphere when it goes flat. [Assume carbon dioxide behaves as an ideal gas under these conditions.]

15 The volume of 1 mole of carbon dioxide was measured at various pressures but at a constant temperature of 285 K. The results obtained were plotted as shown below.



(a) Sketch, on the above axes, a graph to show how the *pV* value should change with pressure for 1 mole of an ideal gas at the same temperature.

(b) Use the graph to calculate the volume of 1 mole of carbon dioxide at a pressure of 10.0×10^5 Pa. Calculate the volume at 285 K that the ideal gas equation predicts for this pressure and comment on the difference between the two values.

16 0.2 mol of radon is stored in a 20.0 dm^3 flask with 0.8 mol of fluorine at 127 °C.

(a) Calculate the total pressure in the flask.

(b) Calculate the partial pressure of radon and fluorine in the flask.

(c) Radon and xenon are in the same group. A scientist hypothesises that radon has similar chemical reaction as xenon. Xenon can react with fluorine according to the following reaction.

$$Xe(g) + F_2(g) \longrightarrow XeF_2(g)$$

The mixture of radon and fluorine in the flask was sparked at 127 °C. What is the final pressure in the flask if the temperature is maintained at 127 °C?

Answer keys for MCQ:

1	2	3	4	5	6	7	8	9	10	11
A	C	C	C	B	D	D	A	A	B	B

12) 495 kPa	13b) M_r = 50.2	14a) 5.18×10^7 Pa	14c) 6.94×10^{-4} m^3
15b) $2.37 \times 10^{-3} m^3$	16a) 166 kPa	16b) P_{Rn} = 33.2 kPa P_{F2} = 133 kPa	16c) 133 kPa