CHEMISTRY
LEVEL 4C
(CHM 415109)

KINETIC THEORY
&
GAS LAWS

THEORY SUMMARY
&
REVISION QUESTIONS

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INTRODUCTION

In simple terms, all matter exists either as gases, liquids or solids. This division of matter into these three states is a simplification as it excludes glasses, plasmas, liquid crystals and a number of other so called ‘states’.
However for our course in chemistry the three main states of solids, liquids and gases will provide adequate coverage.
This particular study unit focuses on gases and the gas phase which is sometimes described as the ‘almost empty’ phase.
We will look at some of the special properties of gases and in particular consider the mathematical laws that are often used to describe gas behaviour.

COMPARISON OF SOLIDS, LIQUIDS & GASES

A comparison of the physical properties of gases, liquids and solids is summarised in the table below.

PROPERTY	GASES	LIQUIDS	SOLIDS
DENSITY	Low density	High density	High density
VOLUME IN CONTAINER	Occupies whole container volume	Fixed volume	Fixed volume
SHAPE	Shape fits the container	Shape changes to fill pa of the container	Shape doesn’t change
COMPRESSIBILITY	Easily compressed	Practically incompressible	Practically incompressible
EASE OF MIXING IN EACH OTHER	Mix rapidly	Mix slowly	Negligible mixing

These properties of gases are best described in terms of gases being made up of very small particles (atoms or molecules) that are on average, separated from each other by large distances compared to the size of the particles.
These particles are in rapid and random motion inside the container and the vast majority of the space in the container is empty space!
The particles of a gas move in straight lines at high speeds on average until they collide with another gas molecule or the containing wall.
The number of particles and their speed of collision with the container wall will determine the gas pressure exerted.

KINETIC THEORY OF GASES

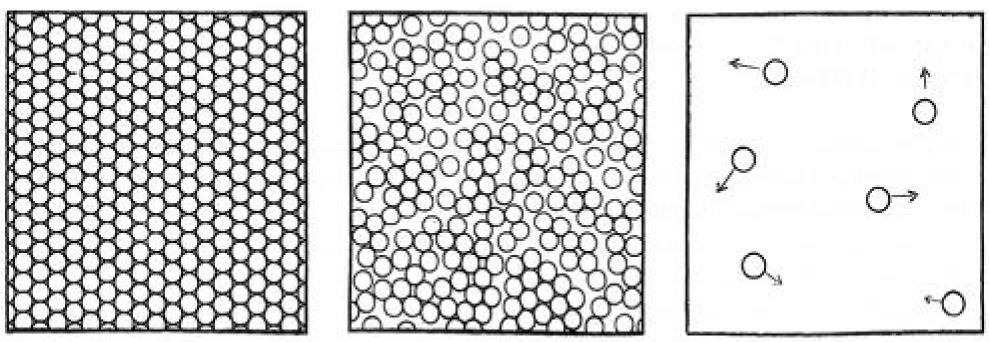
This theory offers a summary of how scientists explain the characteristic properties of gases. Bear in mind that the word ‘kinetic’ infers movement.

(i) Gases comprise tiny particles which are the atoms or molecules of gas.
(ii) The size of the gas particles is incredibly small and the total volume of the actual particles themselves, represents a tiny proportion of the container’s total volume.
(iii) The gas particles are distributed throughout the whole of the container’s volume.
(iv) The gas molecules move in straight lines at very high speeds on average until they collide with other gas molecules or the containing wall.
(v) At any one instant the gas molecules will not all be moving with the same speed. Some will be moving rapidly and some will be momentarily at rest. However, the average speed for all molecules will be very high; (see the details for oxygen below)
(vi) The intermolecular forces (dispersion forces or van der Waal’s forces) between gas molecules are negligible.
(vii) Collisions between gas molecules are perfectly elastic which means that there is no loss of kinetic energy during collisions.
(viii) The average speed of gas molecules is affected by the gas temperature. Raising the gas temperature results in the gas particles moving faster on average and thus having a higher kinetic energy than before.
e.g.

CONSIDER 1.0 L OF OXYGEN AT 25°C AND 1.0 ATMOSPHERE

- There are about 2.5 x 10^{22} molecules of oxygen in this one litre volume.
- The average speed of these oxygen molecules (O_2) at is about 480 m s^{-1} which corresponds to 1730 km/h.
- During each second, each oxygen molecule undergoes about 1 billion collisions with other molecules.
- On average, each oxygen molecule travels about 0.5 microns between collisions. (where 1 micron = 1 micrometre = 1 μm = 1 x 10^{-6} m)

COMPARING THE PACKING



SOLIDS	LIQUIDS	GASES
REGULAR PACKING	RANDOM CLOSE	LARGE SEPARATIONS
VIBRATION ONLY	PACKING	RAPID MOTION

EXPLANATIONS OF GAS PROPERTIES

The Kinetic Theory of Gases provides a good basis for explaining many of the properties of gases that we observe; i.e.

PROPERTY	EXPLANATION
DENSITY	Gases have very low densities because a large proportion of the gas space is unoccupied.
VOLUME & SHAPE	The fast moving gas particles quickly expand to fill the available space.
COMPRESSIBILITY	Gases are easily compressed because the particles can be pushed closer together and now occupy spaces that were between them previously.
EASE OF MIXING	Because of the large spaces between molecules and their high average speeds, the particles can travel relatively large distances before colliding with others. This allows rapid mixing.

GASEOUS ELEMENTS

Of the 100 or so elements, only eleven are gases at room temperature and pressure. These should be known together with the fact that some are monatomic (X) and some are diatomic (X_2).

ELEMENT	SYMBOL	MONATOMIC OR DIATOMIC
hydrogen	H_2	diatomic
helium	He	monatomic
nitrogen	N_2	diatomic
oxygen	O_2	diatomic
fluorine	F_2	diatomic
neon	Ne	monatomic
chlorine	Cl_2	diatomic
argon	Ar	monatomic
krypton	Kr	monatomic
xenon	Xe	monatomic
radon	Rn	monatomic

THE EARTH'S ATMOSPHERE

During the formation of our planet, it retained a tiny envelope of gases which makes up the atmosphere. Its thickness corresponds to about the thickness of water left on the surface of an orange that has been dipped in water.
This fragile band of gases is very delicate and scientists believe that it is under threat from human activities such as the burning of fossil fuels.
The atmosphere can be considered to be roughly 80% nitrogen and 20% oxygen by volume. A more accurate composition of the Earth's atmosphere is shown below.

GAS	% BY VOLUME	SYMBOL/FORMULA
nitrogen	78.1	N_2
oxygen	20.9	O_2
argon	0.9	Ar
carbon dioxide	0.032	CO_2
water vapour	variable	H_2O
hydrogen	trace amounts	H_2
ozone	trace amounts	O_3
methane	trace amounts	CH_4
carbon monoxide	trace amounts	CO
helium	trace amounts	He
neon	trace amounts	Ne
krypton	trace amounts	Kr
xenon	trace amounts	Xe
organics	trace amounts	~
radon	trace amounts	Rn

The gases of the atmosphere can be separated by **FRACTIONAL DITILLATION**. In this process, air is compressed and cooled until it is liquefied. Most of the world's helium gas comes from within underground reserves of crude oil and natural gas. It collected after having been formed in alpha particle emission from radionuclides in the Earth's crust.
By then slowly heating and raising the liquid air's temperature, the individual gases will be boiled off in order of their increasing boiling points (B.P.s).

COMPONENT OF AIR	B.P. (^oC)	B.P. (kelvin)
Helium (He)	-269^oC	4.2 K
Neon (Ne)	-246^oC	27 K
Nitrogen (N_2)	-196^oC	77 K
Argon (Ar)	-186^oC	87 K
Oxygen (O_2)	-183^oC	90 K
Krypton (Kr)	-152^oC	121 K
Xenon (Xe)	-108^oC	165 K

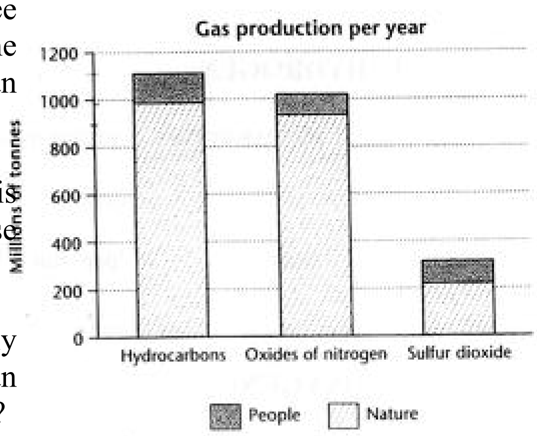
Q1. If liquid air is fractionally distilled, which gas, nitrogen or oxygen will be boiled off first?
(nitrogen)

Q2. By considering the noble gas elements, what is the general relationship between atomic number and boiling point?

ATMOSPHERIC POLLUTION

A pollutant gas is difficult to define although we may regard it as any gas that wouldn't normally be found in the atmosphere. However, some of the so-called 'pollutant' gases are ones that don't come from human activity, but come from natural processes such as volcanic activity.

Consider the bar-graph to the right which shows an estimate of the relative amounts of three 'pollutant' gases produced per year and the proportions due to natural as well as human activity.



Q3. What is the % of each pollutant gas that is caused by human activity according to these estimates?

Q4. If motor vehicles account for approximately 33% of the nitrogen oxides produced by human activity, what is this amount in tonnes per year?

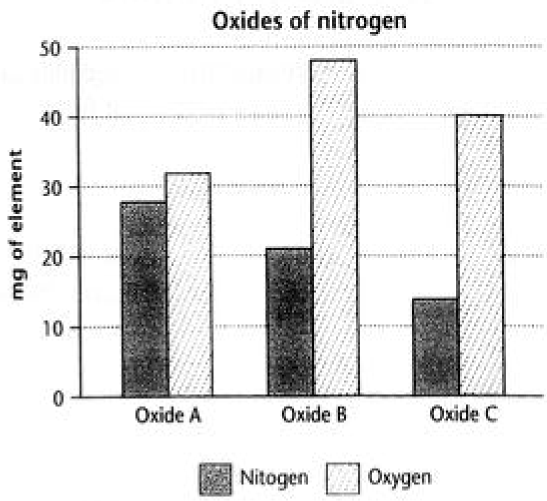
(40 million tonnes)

Q5. What are the possible causes of the natural production of each pollutant gas shown?

Under the high temperature and pressure conditions that exist in diesel and petrol engines, the nitrogen from the air can be made to undergo oxidation forming nitrogen oxides (NO_x) which are a very significant contributor to atmospheric pollution, particularly in their role in forming 'smog'.

Three of the oxides are collected from the polluted air of a busy industrial area. Very small samples of the three oxides were then analysed gravimetrically and the mass of nitrogen to mass of oxygen results shown in the bar-graph to the right.

The oxides were labelled A, B and C



Q6. What mass of nitrogen oxide B was analysed?

(69 mg)

Q7. Use the mass data given to determine the empirical formulae for the three oxides.

(A = NO)

(B = NO_2)

(C = N_2O_5)

GAS PREPARATION & TESTING

In Chemistry 3C you are expected to be aware of some possible methods for preparing and testing gases commonly encountered in the laboratory. The preparations and tests for first three gases below should be known.

1. HYDROGEN:

PREPARATION: reactive metal plus acid, electrolysis of water,.......

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

TEST: 'pop' test $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

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2. OXYGEN:

PREPARATION: decomposing hydrogen peroxide, electrolysis of water, ..

$$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$$

TEST: causes re-ignition of a smouldering splint.

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3. CARBON DIOXIDE:

PREPARATION: acid plus carbonate or hydrogen carbonate, combustion of fuels...

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

TEST: carbon dioxide gas initially turns limewater 'milky' but excess bubbling of carbon dioxide causes the white precipitate of calcium carbonate to redissolve.

$$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow \underbrace{CaCO_{3(s)}}_{\text{MILKY PPT.}} + H_2O_{(l)}$$

$$\underbrace{CO_{2(g)}}_{\text{XS}} + CaCO_{3(s)} + H_2O_{(l)} \rightarrow \underbrace{Ca(HCO_3)_{2(aq)}}_{\text{REDISSOLVES}}$$

4. SULFUR DIOXIDE:

PREPARATION: acid plus sulfite, acid plus thiosulfate,.......

$$Na_2SO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + SO_{2(g)} + H_2O_{(l)}$$

TEST: Sulfur dioxide turns dichromate ions from yellow to green.

$$3SO_{2(g)} + \underset{\textbf{YELLOW}}{Cr_2O_7^{2-}{}_{(aq)}} + 2H^+{}_{(aq)} \rightarrow 3SO_4^{2-}{}_{(aq)} + \underset{\textbf{GREEN}}{2Cr^{3+}{}_{(aq)}} + H_2O_{(l)}$$

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5. HYDROGEN SULFIDE: ('rotten egg' gas)

PREPARATION: acid plus sulfide,.......

$$Na_2S_{(s)} + 2HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2S_{(g)}$$

TEST: rotten egg odour and it turns lead nitrate paper black/silver.

$$H_2S_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow 2HNO_{3(aq)} + PbS_{(s)}$$

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6. CHLORINE:

PREPARATION: oxidation of conc. chloride ions, electrolysis, ...

$$MnO_{2(s)} + 4HCl_{(aq)} \rightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2H_2O_{(l)}$$

TEST: Cl_2 initially turns blue litmus paper red and then bleaches it.
Cl_2 turns sodium bromide solution brown as bromine forms.

$$Cl_{2\,(g)} + 2NaBr_{(aq)} \rightarrow 2NaCl_{(aq)} + Br_{2\,(aq)}$$

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7. AMMONIA:

PREPARATION: ammonium salts plus sodium hydroxide,......

$$NH_4Cl_{(s)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NH_{3(g)} + H_2O_{(l)}$$

TEST: ammonia gas turns red litmus paper blue and ammonia forms a white smoke when mixed with hydrogen chloride gas.

$$NH_{3(g)} + HCl_{(g)} \rightarrow \underset{\textbf{WHITE SOLID}}{NH_4Cl_{(s)}}$$

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GAS PRESSURE MEASUREMENT & UNITS

We now move on to the section of this unit of study where we consider the behaviour of gases and in particular the effect of changing temperatures and pressures on the volume occupied by a certain amount of gas. Before we can start, we need to be familiar with the concept of gas pressure and how it is measured.

GAS PRESSURE

The gas molecules striking the surface of the container's wall exert a force per unit area and this is the basis of 'pressure'; i.e. pressure is defined as $P = (\text{force} \div \text{area})$

The amount of force exerted by the gas molecules on a given area of the containing wall is affected by:

- How fast (on average) the molecules are moving.
- The number of gas molecules in the container.
- The volume of the container

PRESSURE UNITS

There are only three pressure units that you need to be familiar with for this course in chemistry. These are defined below:

(i) **pascal (Pa)** This is the S.I. unit for pressure and is the pressure corresponding to a force of one newton acting over an area of one square metre.

i.e. $1.00 \text{ Pa} = 1.00 \text{ N m}^{-2}$

A pascal is such a small pressure that most of the time we need to be working in kilopascals or even megapascals. ($1 \text{ MPa} = 1000 \text{ kPa} = 10^6 \text{ Pa}$)

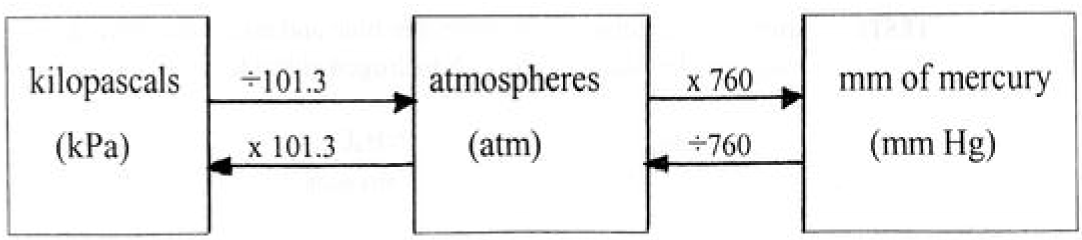
(ii) **atmosphere (atm)** This is the pressure that occurs on average at sea level on Earth and is due to the weight of air above us. Standard atmosphere pressure is 101.3 kPa.

(iii) **millimetres of mercury (mm Hg)** Pressure can be measured in terms of a height of a column of liquid where the height is inversely proportional to the liquid's density.

The Earth's atmosphere is able to support a column of mercury of height about ¾ of a metre. This varies from day to day depending upon high and low pressure weather patterns. However, 'standard' atmospheric pressure is 1.000 atmosphere = 760.0 mm Hg

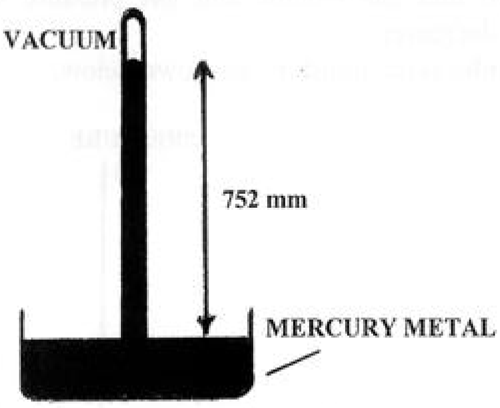
CONVERSION OF PRESSURE UNITS

$$1.000 \text{ atm} = 101.3 \text{ kPa} = 760.0 \text{ mm Hg}$$

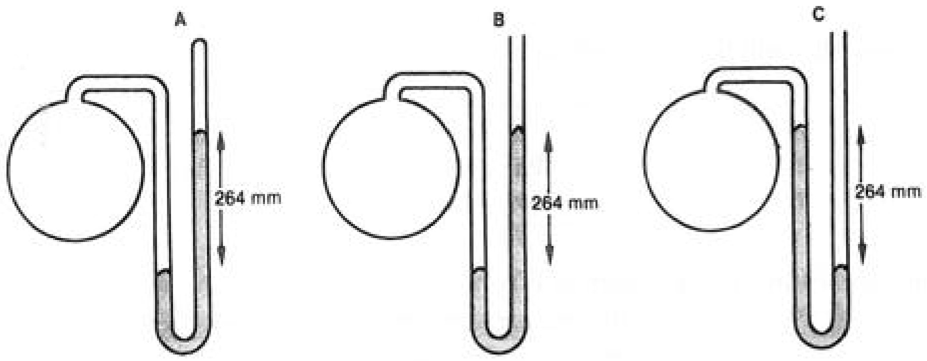


MERCURY BAROMETERS & MANOMETERS

The density of liquid mercury metal makes it ideal for measuring pressures of gases. The technique that we use, often utilises "U" tubes filled with mercury metal and by careful measurement of vertical heights we can translate these heights to give us pressure. The diagrams below will help you to understand the technique.



As shown in the diagram above, the gas pressure measurements occurred on a day when the prevailing atmospheric pressure was 752 mm Hg.
Now assuming that the atmospheric pressure was 752 mm Hg, let us now consider three different gas systems where a mercury manometer has been used.
Remember that if the manometer has an ***open*** end as occurs in cases (ii) and (iii) below, it means that the gas pressure must be compared to the atmospheric pressure as being either greater or less than atmospheric. (The diagrams are ***not*** to scale!)



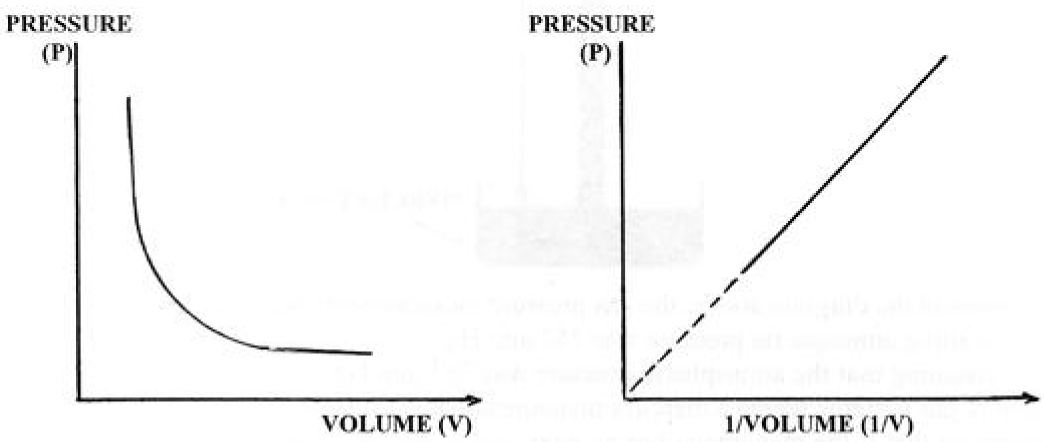
(In each of the manometers above, the liquid present in the 'U' tubes is mercury)

Q8. Calculate the pressure in the 3 gas chambers above in units of (i) mm Hg, (ii) kPa.
(Remember that the atmospheric pressure was 752 mm Hg)

ANS.	**A: 264 mm Hg**	**B: 1016 mm Hg**	**C: 488 mm Hg**
	A: 35.2 kPa	**B: 135 kPa**	**C: 65.0 kPa**

BOYLE'S LAW (THEORY)

In the mid 1600s, Robert Boyle investigated the effects of pressure on gas volumes.
His investigations were carried out using a fixed amount of gas at a constant temperature.
His results revealed that gas volume and gas pressure vary inversely; i.e. as one increases, the other decreases.
His results gave graphical relationships as shown below.



The second graph shown above indicates a linear relationship between pressure and reciprocal volume; i.e.

$P \propto 1/V$ (provided T and mass of gas = const.)
$P = (1/V) \times \text{constant}$
$P \times V = \text{constant}$

This translates into the more commonly encountered equation for Boyle's Law; i.e.

$$\boxed{\mathbf{P_1V_1 = P_2V_2}}$$

In this mathematical expression of Boyle's Law,
P_1 = initial gas pressure and P_2 = final gas pressure
V_1 = initial gas volume and V_2 = final gas volume

Boyle's Law states that "for a fixed amount of gas at a constant temperature, the volume of the gas varies inversely with the pressure."

e.g. Consider 10 L of gas at 1.0 atmosphere pressure. Halving of the gas volume to 5.0 L would mean that these same number of molecules are now twice as close and thus strike the container wall twice as often. This means the pressure will have doubled to 2.0 atm.
This fits the Boyle's Law expression;
$P_1V_1 = P_2V_2$ because $1.0 \times 10 = 2.0 \times 5.0$

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QUESTIONS INVOLVING BOYLE'S LAW

Q9. If the pressure on a given amount of gas is made three times greater, what happens to its volume? (Assume the temperature is constant) **(1/3rd original)**

Q10. A given sample of neon gas occupies a volume of 12.0 L at a pressure of 150 kPa. What volume will the $Ne_{(g)}$ occupy if the pressure is increased to 175 kPa and the temperature held constant? **(10.3 L)**

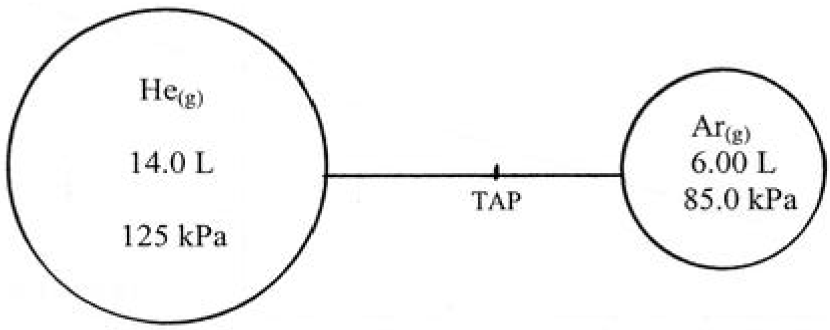
Q11. A scuba cylinder has an internal volume of 11.2 L and contains medical air at a pressure of 22.4 MPa. What volume will this air occupy at a pressure of one atmosphere (101.3 kPa) and fixed temperature? **(2480 L)**

Q12. The volume of a sample of nitrogen gas is 355 mL at a pressure of 0.875 atm. The gas is allowed to expand at constant temperature to a volume of 425 mL. What is the new pressure? **(0.731 atm)**

Q13. If 25.0 L of helium gas at 1.35 atm is compressed to a pressure of 175 kPa, what will be the final volume? (T constant) **(19.5 L)**

Q14. A car tyre has an internal volume of 18 L and is pumped to a total pressure of 303 kPa on a day when the atmospheric pressure is 101 kPa. What volume of air is needed to pump up the tyre? (Remember that there's already air in a "flat" tyre and assume that the volume of the flat tyre is also 18 L) **(36 L)**

Q15. The two gas containers shown below are connected by a narrow tube of negligible volume including a closed tap.



What will be the total pressure in the system above when the tap is opened and the gases mix? Assume that the temperature is held constant throughout. **(113 kPa)**

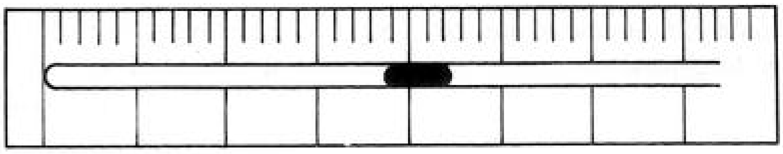
Q16. The following data were obtained for a fixed sample of nitrogen gas at a constant temperature of 25°C.

PRESSURE (mm Hg)	**376**	**522**	**608**	**742**	**871**	**909**	**1070**
VOLUME (mL)	**225**	**162**	**139**	**114**	**97**	**93**	**79**

Show both numerically and graphically that these data obey Boyle's Law.

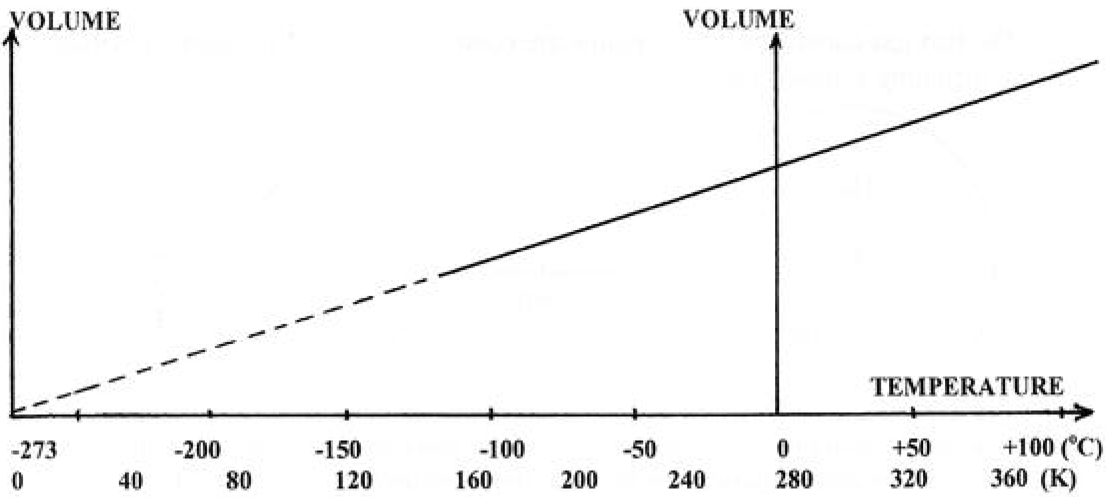
CHARLES' LAW (THEORY)

At constant pressure gases expand when heated. However, raising the temperature of a fixed mass of gas from 10^{o}C to 20^{o}C certainly doesn't double the volume. The volume only increases by about 3%.
Charles' Law enables us establish exactly what volume changes occur when gases are heated or cooled.
The apparatus below shows a sample of dry nitrogen gas in a capillary tube with one end open and the other closed. The tube is attached to a calibrated scale which enables the gas volume to be calculated.



APPARATUS USED FOR TESTING CHARLES' LAW

The gas sample was then heated and the volume recorded. These data, when plotted, give a straight line graph which is shown below.



Charles found that other gases gave similar linear results with the data always giving a straight line that when extrapolated back to zero volume, intersected the temperature axis at −273.15^{o}C. This temperature point represents a theoretical value where the gas volume would become zero. To three significant figures we use −273^{o}C.

This temperature is referred to as **ABSOLUTE ZERO** and becomes the zero point for a new and extremely useful temperature scale called the absolute temperature scale.
The absolute temperature scale is in units of kelvin (K) where 1 kelvin unit is the same as 1 degree celsius unit. The difference is that the two temperature scales have different zero points.

KELVIN TEMPERATURE SCALE

To convert from celsius temperatures to kelvin temperatures we use the relationship:

Temperature in kelvin (K) = (temperature in ^{o}C + 273)

The reason this new temperature scale is so important is that it allows us to now see a ***direct*** relationship between the temperature and the volume of a gas. The graph on the previous page passes through the origin providing we use the kelvin temperature scale; i.e.

$V \propto T$ (provided T is in kelvin and P = const.)

$V = \text{constant} \times T$

$V/T = \text{constant}$

This translates into the more commonly encountered equation for Charles' Law; i.e.

$$\mathbf{V_1/T_1 = V_2/T_2}$$

In this mathematical expression of Charles' Law,

V_1 = initial gas volume V_2 = final gas volume

T_1 = initial gas temperature in kelvin T_2 = final gas temperature in kelvin

Charles' Law states that "for a fixed amount of gas at a constant pressure, the volume of the gas varies directly with the *absolute* temperature."

Charles' Law is equally applicable to considerations of the pressure change of a gas when the temperature changes. In this case, the volume of gas is held constant and the pressure is measured at various temperatures.

In a very similar way, the relationship $\mathbf{P_1/T_1 = P_2/T_2}$ is derived.

EXAMPLE: A sample of argon gas occupies a volume of 176 mL at a temperature of 16^{o}C. What volume will it occupy if the temperature is increased to 45^{o}C with the pressure remaining unaltered?

ANS. V_1 = initial gas volume = 176 mL

V_2 = final gas volume = ? mL

T_1 = initial gas temperature in kelvin = (16 + 273) = 289 K

T_2 = final gas temperature in kelvin = (45 + 273) = 318 K

Substitute into $V_1/T_1 = V_2/T_2$ gives $(176/289) = (V_2/318)$

i.e. $V_2 = \{(176 \times 318)/289\}$ mL = 194 mL

i.e. Final volume of argon = 194 mL.

COMBINED GAS EQUATION

Boyle's Law and Charles' Law can be combined to produce the useful mathematical relationship:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \qquad \text{(for a constant amount of gas)}$$

Where '**1**' refers to initial conditions and '**2**' refers to final conditions.

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Q17. Some helium gas is at a temperature of 10°C. To what temperature (in celsius) must it be heated in order to double its volume, assuming constant pressure?

(293°C)

Q18. A car tyre is pumped to a total pressure of 285 kPa on a cold morning when the temperature is 4°C. What will be the tyre pressure when having travelled for some time, road friction has heated the tyre to 45°C? (assume no volume change) **(327 kPa)**

Q19. A sample of argon gas occupies a volume of 250 mL at 20°C and 120 kPa. What volume will it occupy at 40°C and 120 kPa? **(267 mL)**

Q20. A sample of methane gas occupies a volume of 174 mL at 35°C and 125 kPa. What volume will it occupy at 100°C and 108 kPa? **(244 mL)**

Q21. A meteorological balloon of initial volume 40.0 L is inflated at ground level with hydrogen gas at a pressure of 1.00 atm and a temperature of 25°C. What will be the balloon's volume when it reaches an altitude where the pressure is 0.300 atm and the temperature is −20°C? **(113 L)**

Q22. A given sample of oxygen gas has an initial volume of 5.00 L at 20.0°C and 111 kPa. After changes have occurred, the new volume is 5.65 L at a pressure of 101kPa. What is the final temperature of the oxygen gas? **(28.3 °C)**

Q23. Fill in the missing numbers corresponding to letters a, b, c,…… etc in the following table of data relating to fixed masses of various gases that have been subjected to changes in temperature or pressure or volume.

P_1	V_1	$Temp_1$	P_2	V_2	$Temp_2$
101.3 kPa	3.50 L	290 K	48.6 kPa	a?	430 K
1.60 atm	462 mL	293 K	b?	725 mL	388 K
1.00 atm	250 mL	25°C	106.6 kPa	0.600 L	c?
442 mm Hg	0.370 L	22°C	93.0 kPa	450 mL	d?
67 kPa	2.37 L	461°C	e?	788 mL	512 K

(a = 10.8 L) **(b = 1.35 atm)** **(c = 753 K or 480°C)**
(d = 566 K or 293°C) **(e = 141 kPa)**

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PARTIAL PRESSURES

THEORY: In a mixture of gases A, B, C,..... the total pressure is the sum of the partial pressures of the component gases and the partial pressure of each component is directly proportional to the mole fraction of that component. This is called Dalton's Law.

i.e. $\mathbf{P_{TOT} = p_A + p_B + p_C + \ldots\ldots}$ where: p_A = partial pressure of gas A

P_{TOT} = total pressure

$\mathbf{p_A = \{n_A/n_{TOT}\} \times P_{TOT}}$ n_A = moles of gas A = $n(A_{(g)})$

n_{TOT} = total moles of gas

n_A/n_{TOT} = mole fraction of A

Q24. A mixture contains 2.00 moles of oxygen gas and 3.00 moles of nitrogen gas. What is the mole fraction of $O_{2(g)}$? **(0.400 or 40.0%)**

Q25. If the mixture of gases in Q1.(above) is at a total pressure of 125 kPa, what is the partial pressure of the oxygen? **(50.0 kPa)**

Q26. A gas mixture contains 0.600 mole of neon, 0.500 mole of argon and 0.900 mole of helium all at a total pressure of 780 mmHg. What is the partial pressure exerted by the helium gas ? **(351 mmHg)**

Q27. A container holds a mixture of krypton gas $Kr_{(g)}$ at a partial pressure of 50.0 kPa and helium gas $He_{(g)}$ at a partial pressure of 100.0 kPa. For this system find:

(i) the total pressure. **(150 kPa)**

(ii) the mole fraction of helium. **(0.667 or 66.7%)**

Q28. A mixture of gases contains equal ***masses*** of oxygen (O_2) and methane(CH_4). The total pressure of the mixture is 3.30 atm. For this system find:

(i) the mole fraction of oxygen. **(0.333 or 33.3%)**

(ii) the partial pressure of the methane gas. **(2.20 atm)**

Q29. Normal atmospheric air contains 21.0% oxygen by volume. What is the partial pressure of oxygen in pure air when the total pressure is 104 kPa? **(21.8 kPa)**

Q30. Three gases A, B and C form a mixture where the partial pressures are:
p_A = 55.0 kPa, p_B = 65.0 kPa and p_C = 75.0 kPa.

(i) What is the total pressure of the mixture? **(195 kPa)**

(ii) What is the mole fraction of gas A in the mixture? **(0.282 or 28.2%)**

Gas component B is now removed from the mixture, leaving only gases A and C.

(iii) What is the total pressure of the gas mixture now? **(130 kPa)**

(iv) What is the mole fraction of gas A in the mixture now? **(0.423 or 42.3%)**

(v) What is the partial pressure of gas A now? **(55.0 kPa–unaltered!)**

Q31. A 12.0 L container holds argon gas at a pressure of 115 kPa. This container is connected by a narrow tube incorporating a closed tap to a container of volume 22.0 L holding nitrogen gas at a pressure of 106 kPa. The tap is opened and the gases mix. The temperature is unaltered. What is the total pressure of the final gas mixture?

(109 kPa)

GENERAL GAS EQUATION THEORY

1. The General Gas Equation is used extensively in gas calculations to find the inter-relation between pressure, volume, temperature and amount of gas.

2. The General Gas Equation assumes that the gases concerned are "perfect" i.e. behave "ideally" which implies that the gas molecules exert no intermolecular forces (van der Waal's or dispersion forces) on one another.

3. Real gases behave most closely to being 'ideal' when the pressures are generally low and the temperatures are high.

4. At S.L.C. (1.00 atm and 298 K) one mole of gas occupies a volume of 24.5 L

5. At S.T.P. (1.00 atm and 273 K) one mole of gas occupies a volume of 22.4 L

i.e. applying Boyle's Law, $P \times V = \text{constant}$, gives, for 1.00 mole of gas

$$P \times V = 22.4 \quad \text{L.atm} \qquad \text{at } 0^{o}\text{C} \quad (273\text{ K})$$

so, for 'n' moles of gas $\quad P \times V = 22.4 \times n \quad \text{L.atm} \quad \text{at } 273\text{ K}$

Thus for 'n' moles of gas at T kelvin, applying Charles' Law,

$$P \times V = 22.4 \times n \times (T/273) \quad \text{L.atm}$$

or: $\quad P \times V = n \times \dfrac{22.4}{273} \times T \qquad (22.4/273) = 0.0821$

i.e. $\quad P \times V = n \times R \times T \qquad$ where R =ideal gas constant

$$\boxed{\mathbf{P.V = n.R.T}}$$

NOTE: The value of the ideal gas constant R is dependent upon the units used.

Pressure units (P)	**Volume units (V)**	**Value of R**
kilopascals (kPa)	litres (L)	8.31
atmospheres (atm)	litres (L)	0.0821
mm of mercury (mm Hg)	litres (L)	62.4

NOTE: For all calculations using the General Gas equation, ensure that the temperature is in kelvin!

DIFFERENT FORMS OF THE GENERAL GAS EQUATION

For a single gas:	$PV = nRT$	R = ideal gas constant mass of gas in grams T = temperature in kelvin
	$PV = \frac{(\text{mass})}{M_r}RT$	M_r = relative molecular mass
For a mixture of gases:	$p_AV = n_ART$	n_A = moles of gas A i.e. $n(A_{(g)})$ p_A = partial pressure of gas A
or:	$P_{TOT}V = n_{TOT}RT$	n_{TOT} = total moles of gas P_{TOT} = total gas pressure

GENERAL GAS EQUATION PROBLEMS

Q32. What is the pressure exerted by 18.0 g of carbon dioxide gas in 20.0 L container at 55°C?

(55.8 kPa)

Q33. What is the volume occupied by 40.0 g of nitrogen gas (N_2) at a pressure of 1.35 atm and a temperature of 70.0°C?

(29.8 L)

Q34. What mass of ethane gas (C_2H_6) occupies a volume of 90.0 L at a pressure of 88.0 kPa and a temperature of 30°C?

(94.4 g)

Q35. To what temperature (in °C) must 38.0 g of argon gas (Ar) be heated in order to exert a pressure of 110 kPa in a container of volume 22.0 L?

(33.5°C)

Q36. A 5.60 g sample of the gaseous ***element*** X_2 at 25°C occupies a volume of 3.51 L at a pressure of 104 kPa. Find:

(i) $M_r(X_2)$ **(38.0)**
(ii) $A_r(X)$ **(19.0)**
(iii) the likely identity of element X. **(fluorine)**

Q37. Assuming that air behaves as a single gas with M_r = 28.8, find the mass of air at S.L.C. in a laboratory with dimensions of 12.0 x 7.00 x 3.00 m^3. (1 m^3 = 1000 L)

(297 kg)

Q38. What is the density of propane gas (C_3H_8) at 100°C and a pressure of 767 mm Hg? Give your answer in units of g L^{-1}. **(1.45 g L^{-1})**

Q39. Consider a container of volume 20.0 L at 25.0°C holding a mixture of methane gas $CH_{4(g)}$ at a partial pressure of 50.0 kPa and helium gas $He_{(g)}$ at a partial pressure of 100.0 kPa.
For this system find:

(i) the mass of methane gas. **(6.46 g)**

(ii) the mass of helium gas. **(3.23 g)**

(iii) the density of the mixture in units of g L^{-1}. **(0.485 g L^{-1})**

(iv) the total pressure. **(150 kPa)**

(v) the mole fraction of methane. **(0.333 or 33.3%)**

(vi) the mole fraction of helium. **(0.667 or 66.7%)**

Q40. A gas mixture contains 11.2 g of $N_{2(g)}$, 9.60 g of $CH_{4(g)}$ and 64.0 g of $O_{2(g)}$ in a 80.0 L flask at 27°C .
Find:

(i) the partial pressure of each component (units of kPa).
(pN_2 = 12.5 kPa)
(pCH_4 = 18.7 kPa)
(pO_2 = 62.3 kPa)

(ii) the total pressure in kPa. **(93.5 kPa)**

(iii) the mole fraction of methane in the mixture. **(0.200 or 20.0%)**

(iv) the partial pressure of $N_{2(g)}$ if all the $O_{2(g)}$ is removed. **(12.5 kPa)**

Q41. Assuming that air is 80.0% N_2 molecules and 20.0% O_2 molecules, find the mass of

(i) oxygen in 1.00 L of air at 1.00 atm and 20.0°C. **(0.266 g)**

(ii) nitrogen in 1.00 L of air at 1.00 atm and 20.0°C. **(0.931 g)**

Q42. Consider the chemical reaction:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

By identifying the **limiting** reagent, find the volume of carbon dioxide gas (CO_2) produced at 101.3 kPa and 25°C when 10.0 g of calcium carbonate ($CaCO_3$) is reacted with 50.0 mL of 2.00 mol L^{-1} $HCl_{(aq)}$? **(1.23 L)**

Q43. Consider the balanced equation:

$$3Cu_{(s)} + 8HNO_{3(aq)} \rightarrow 3Cu(NO_3)_{2(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

If 5.00 g of copper metal is reacted with excess nitric acid in accordance with the equation given, what will be the volume of nitric oxide gas (NO) formed at 25°C and 108 kPa pressure? **(1.20 L)**

Q44. A gaseous compound X has a density of 1.86 g L^{-1} at 112°C and 1.12 atmosphere pressure. Use these data to find M_r(X). **(52.5)**

GENERAL REVISION QUESTIONS ON GASES

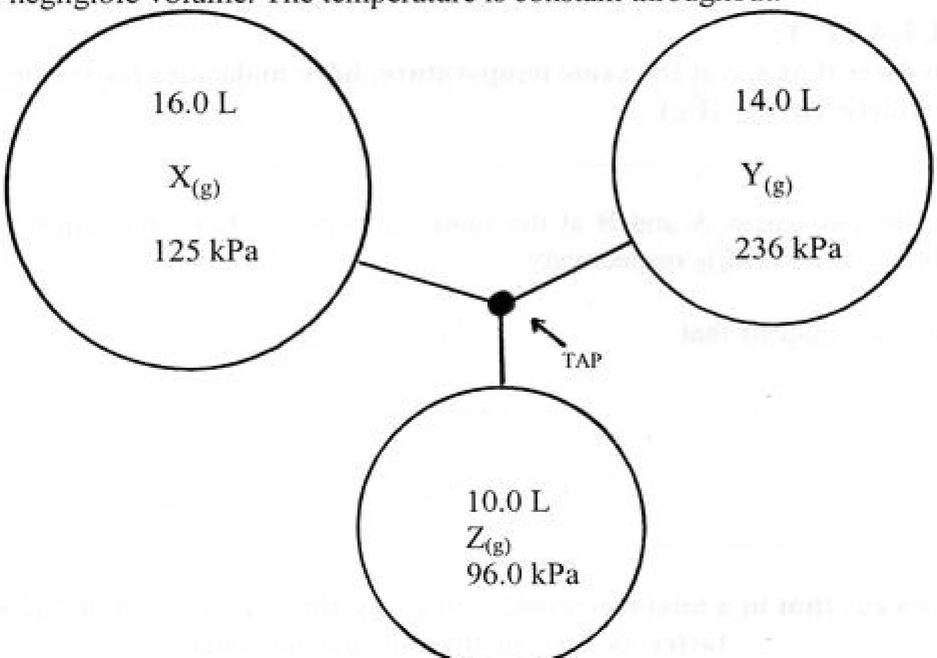
Q45. A sample of neon gas occupies 5.00 L at 27°C and 108 kPa. What volume will it occupy at 10°C and 95.0 kPa?

(5.36 L)

Q46. What is the volume of 100.0 g of carbon dioxide gas at 17°C and 735 mm Hg?

(56.0 L)

Q47. Consider the three gas containers shown below. They are all interconnected by tubes of negligible volume. The temperature is constant throughout.



(i) What are the partial pressures of each component gas in the system if the tap is opened and the gases are allowed to mix? Assume that the temperature is constant throughout and that gases X, Y and Z do not undergo any reactions.

(pX = 50.0 kPa)
(pY = 82.6 kPa)
(pZ = 24.0 kPa)

(ii) What will be the total pressure after the mixing has occurred?

(156.6 kPa)

Q48. A mixture of gases initially contains 14.0 g of $N_{2(g)}$, 9.60 g of $CH_{4(g)}$ and 54.4 g of $O_{2(g)}$ in a 70.0 L flask at 20°C .

(i) Find: (a) the total pressure (use units of kPa). **(97.4 kPa)**
(b) the mole fraction of methane in the mixture. **(0.214 or 21.4%)**

This gas mixture is now sparked and complete combustion of methane occurs until one of the reactants (either methane or oxygen) is used up. The nitrogen does not undergo any reaction. After the reaction, the temperature is restored to the original 20°C and the water vapour from the reaction condenses to form liquid water.

(ii) Assuming the liquid water occupies negligible volume, what is the final pressure in the 70.0 L flask at 20°C? **(55.7 kPa)**

(iii) Will the partial pressure of the nitrogen have changed? **(unaltered)**

KINETIC THEORY FOR GASES
QUANTITATIVE ASPECTS

The properties of gases that we have encountered lead us to 2 very important postulates about gas behaviour and the average kinetic energy ($\bar{E}_K$) of their molecules.
Remember that $E_K = ½mv^2$. We use $\underline{v}$ for average molecular speed and $\bar{E}_K$ for average kinetic energy.

POSTULATE 1:
Any two gases that are at the same temperature, have molecules possessing the same average kinetic energy ($\bar{E}_K$).

Consider the two gases A and B at the ***same*** temperature. Let their relative molecular masses be $(M_r)_A$ and $(M_r)_B$ respectively.

This postulate suggests that $(\bar{E}_K)_A = (\bar{E}_K)_B$

Thus: $(½M_r\underline{v}^2)_A = (½M_r\underline{v}^2)_B$

$$(M_r)_A/(M_r)_B = \underline{v}^2_B/\underline{v}^2_A$$

This means that in a mixture of different gases, the lighter gas molecules travel faster on average than the heavier ones.

Q49. Compare the average speeds of helium gas atoms (He) and methane molecules (CH_4) in a mixture of the two gases. **($\underline{V}_{He}$= 2 x $\underline{v}_{CH4}$)**

Q50. Compare the average speeds of hydrogen molecules (H_2) and oxygen molecules (O_2) in a mixture of the two gases. **($\underline{V}_{H2}$= 4 x $\underline{v}_{O2}$)**

Q51. If the average speed of oxygen molecules (O_2) at 25°C is 481 ms^{-1}, what is the average speed of:
- (i) nitrogen molecules (N_2) at 25°C? **(514 ms^{-1})**
- (ii) hydrogen molecules (H_2) at 25°C? **(1924 ms^{-1})**

Q52. Why do hydrogen filled balloons deflate much more quickly than the equivalent balloons filled with air?

Q53. Given that gas X molecules have an average speed that is **six** times greater than the average speed of gas Y molecules:
- (i) Compare their relative molecular masses. **(M_r(Y) = 36 x M_r(X))**
- (ii) If gas Y is pentane, C_5H_{12}, what is the likely identity of gas X? **(X = hydrogen)**

POSTULATE 2:
The average kinetic energy ($\bar{E}_K$) of gas molecules is directly proportional to the absolute temperature of the gas.

Consider a gas being heated from T_1 kelvin to T_2 kelvin. This postulate suggests that the average kinetic energy of the gas molecules will change from $(\bar{E}_K)_1$ to the new value of $(\bar{E}_K)_2$ where:
$(\bar{E}_K)_1/(\bar{E}_K)_2 \quad = \ T_1/T_2$

$$(\tfrac{1}{2}m\underline{v}_1^{\,2})/(\tfrac{1}{2}m\underline{v}_2^{\,2}) \ = \ T_1/T_2$$

As the molecular mass of the gas doesn't alter with temperature we cancel the (½m) terms and get:

$$\boxed{(\underline{v}_1^{\,2})/(\underline{v}_2^{\,2}) \ = \ T_1/T_2}$$

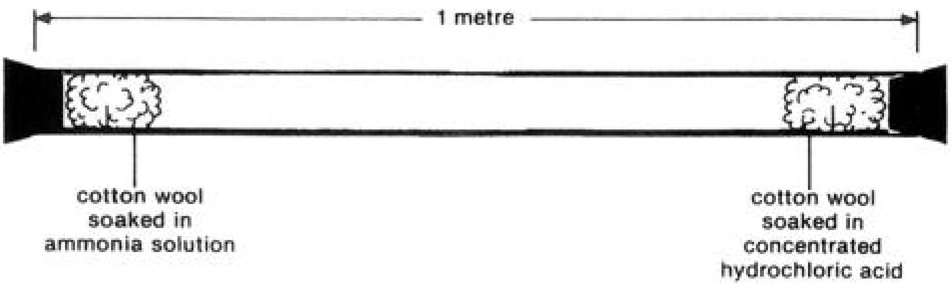
Q54. If the average speed of nitrogen molecules (N_2) at 25°C is 514 ms^{-1}, what is their average speed at 50°C? **(535 ms^{-1})**
Q55. If the average speed of oxygen molecules (O_2) at 25°C is 481 ms^{-1}, what is the average speed of:
- (i) oxygen molecules (O_2) at 125°C? **(556 ms^{-1})**
- (ii) nitrogen molecules (N_2) at 125°C? **(594 ms^{-1})**

DIFFUSION OF GASES

Diffusion is the mixing of substances in each other without there being agitation or shaking.
Q56. Gases diffuse into other gases at much greater speeds than liquids diffuse into other liquids. Explain this observation.
Q57. The diffusion of solids into other solids is almost non-existent. Why?

The rate of diffusion of a gas is directly related to its average molecular speed. Thus, under the same temperature conditions, gases with lower M_r values diffuse more quickly than those with higher M_r values. The rate of diffusion ∞ $(1/\sqrt{M_r})$

Q58. Consider the following experiment involving the equipment as shown below. The cotton wool balls release NH_3 gas at the left end and HCl gas at the right end.



Ammonia gas (NH_3) and hydrogen chloride gas (HCl) react to form a white solid NH_4Cl.
As the two gases diffuse towards each other in the tube above, predict where the formation of the white solid (NH_4Cl) will appear.

CHEMISTRY (LEVEL 4C)
GASES - TEST
CRITERIA 7 & 8

TOTAL = 30 marks

Q1. A sample of argon gas occupies 112 mL at 25°C and 118 kPa pressure. What volume will it occupy at 65°C and 78.0 kPa pressure?

(5 marks)

Q2. What volume is occupied by 53.0 g of chlorine gas ($Cl_{2(g)}$) at 37°C and a pressure of 114 kPa?

(5 marks)

Q3. A pure gas X has a density of 1.86 g L^{-1} at 112°C and 1.12 atm pressure. Use these data to find $M_r(X)$.

(5 marks)

Q4. A mixture of gases contains 27.2 g of ammonia gas $NH_{3(g)}$ and 25.6 g of oxygen gas $O_{2(g)}$ in a 45.0 L flask at 27°C .

Find (a) the total pressure (use units of kPa).

(b) the mole fraction of ammonia in the mixture.

(c) the partial pressure of ammonia in the mixture.

(7 marks)

Q5. Consider the balanced equation:

$$3Cu_{(s)} + 8HNO_{3(aq)} \rightarrow 3Cu(NO_3)_{2(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

If 5.00 g of copper metal is reacted with excess nitric acid in accordance with the equation given, what will be the volume of nitric oxide gas (NO) formed at 25°C and 108 kPa pressure?

(8 marks)

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TEST ANSWERS:

NAME ANSWERS

GASES - TEST

CRITERION 8.

TOTAL = 30 marks

Q1. A sample of argon gas occupies 112 mL at 25°C and 118 kPa pressure. What volume will it occupy at 65°C and 78.0 kPa pressure?

$$\left.\begin{array}{l} V_1 = 112\,mL \\ P_1 = 118\,kPa \\ T_1 = 298\,K \end{array}\right\} \longrightarrow \left.\begin{array}{l} V_2 = ?\,mL \\ P_2 = 78.0\,kPa \\ T_2 = 338\,K \end{array}\right\}$$

$$\therefore V_2 = \left(112 \times \frac{118}{78.0} \times \frac{338}{298}\right) mL$$

$$= 192\,mL$$

$\therefore$ Final volume of $Ar_{(g)} = 192\,mL$

(5 marks)

Q2. What volume is occupied by 53.0 g of chlorine gas ($Cl_{2(g)}$) at 37°C and a pressure of 114 kPa?

$$\left.\begin{array}{l} P = 114\,kPa \\ V = ?\,L \\ n = \frac{53.0}{71.0} = 0.746\,mol \\ R = 8.31 \\ T = 310\,K \end{array}\right\} \quad V = \frac{nRT}{P} = \left(\frac{0.746 \times 8.31 \times 310}{114}\right) L$$

$$= 16.9\,L$$

$\therefore$ Volume of $Cl_{2(g)} = 16.9\,L$

(5 marks)

Q3. A pure gas X has a density of 1.86 g L^{-1} at 112°C and 1.12 atm pressure. Use these data to find $M_r(X)$.

Consider 1.00L

$$\left.\begin{array}{l} P = 1.12\,atm \\ V = 1.00\,L \\ n = \left(\frac{1.86}{M_r}\right) mol \\ R = 0.0821 \\ T = 385\,K \end{array}\right\} \quad \begin{array}{l} P.V = n.R.T \\ 1.12 \times 1.00 = \frac{1.86}{M_r} \times 0.0821 \times 385 \\ \therefore M_r = \left(\frac{1.86 \times 0.0821 \times 385}{1.12 \times 1.00}\right) \end{array}$$

$$= 52.5$$

$$\therefore \quad M_r(X) = 52.5$$

Q4. A mixture of gases contains 27.2 g of ammonia gas $NH_{3(g)}$ and 25.6 g of oxygen gas $O_{2(g)}$ in a 45.0 L flask at 27°C .

Find (a) the total pressure (use units of kPa).

(b) the mole fraction of ammonia in the mixture.

(c) the partial pressure of ammonia in the mixture.

$n(NH_3) = \left(\frac{27\cdot2}{17\cdot0}\right) = 1\cdot60\text{ mol}$

$n(O_2) = \left(\frac{25\cdot6}{32\cdot0}\right) = 0\cdot80\text{ mol}$

$\therefore n_{TOT} = 2\cdot40\text{ mol}$

Thus $PV = nRT$

(a) $\Rightarrow P_{TOT} = \frac{n_T \times R \times T}{V}$

$= \left(\frac{2\cdot40 \times 8\cdot31 \times 300}{45\cdot0}\right)\text{ kPa}$

$\therefore$ Total press. = 133 kPa

(b) mole fractⁿ $NH_3 = \frac{1\cdot60}{2\cdot40}$

$= 0\cdot667$

(c) $p_{NH_3} = \text{mol fr}^n \times P_{TOT}$

$= 0\cdot667 \times 133$

$p_{NH_3} = 88\cdot6\text{ kPa}$

(7 marks)

Q5. Consider the balanced equation:

$$3Cu_{(s)} + 8HNO_{3(aq)} \longrightarrow 3Cu(NO_3)_{2(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

If 5.00 g of copper metal is reacted with excess nitric acid in accordance with the equation given, what will be the volume of nitric oxide gas (NO) formed at 25°C and 108 kPa pressure?

$n(Cu)$ reacting $= \left(\frac{5\cdot00}{63\cdot5}\right)$

$= 0\cdot0787$ mole

$\therefore n(NO)$ produced $= \left(0\cdot0787 \times \frac{2}{3}\right)$ mole

$= 0\cdot0525$ mole

$\therefore V = \frac{n \cdot R \cdot T}{P} = \left(\frac{0\cdot0525 \times 8\cdot31 \times 298}{108}\right) L$

$= 1\cdot20 L$

$\therefore$ volume of $NO_{(g)}$ is $1\cdot20 L$