



ABOUT KEY SYMBOLS

TO UNLOCK THE

OVERVIEW

OF THE TOPIC

TO SEARCH FOR

ABOUT SCIENCE (CHEMISTRY) KEY-POINTS

Aims to prepare students for the GCE 'O' Level Examinations. This book covers the syllabus for Science (Chemistry).

Each chapter of the book consists of the following:

Learning Objectives

Learning objectives of each chapter based on the syllabus are given in this section. Students should make sure that they are able to meet all the basic requirements of the objectives before they sit for the examinations.

Useful Websites

Some useful websites are listed for students' reference and further research. These websites can help students understand the concepts of the chapters better.

Overview Of The Topic

This table gives students a concept map of the topic. It shows students the structural links between the sub-areas of each topic.

Key Points

This part features concise revision notes to help students concentrate on more important areas. All essential key points are presented to help students to consolidate their knowledge learnt in class.

Stop & Think (Sample Questions With Worked Solutions)

For each topic, questions have been consolidated to check and evaluate students' understanding. This ensures that doubts are better clarified before they proceed to the next section.

Examlips

ExamTips in this book summarise all of students' common mistakes, weaknesses and misconceptions shown in past examinations. Some important reminders for students are also listed in this section.

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CS CamScanner



NO.	1	EXPERIMENTAL TECHNIQUES
NO.	2	THE PARTICULATE NATURE OF MATTER
NO.	3	ATOMIC STRUCTURE AND CHEMICAL BONDING
NO.	4	STOICHIOMETRY AND THE MOLE CONCEPT
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EXPERIMENTAL TECHNIQUES

LEARNING OBJECTIVES

Candidates should be able to:

1.1 Experimental Design

- name appropriate apparatus for the measurement of time, temperature, mass and volume, including burettes, pipettes, measuring cylinders and gas syringes
- suggest suitable apparatus, given relevant information, for a variety of simple experiments, including the collection of gases and the measurement of rates of reaction

1.2 Methods of Purification and Analysis

- describe methods of purification by the use of a suitable solvent, filtration, crystallisation and distillation (including the description but not use of fractional distillation – refer to the fractional distillation of crude oil and fermented liquor)
- 🖾 suggest suitable methods of purification, given information about the substances involved
- describe paper chromatography and interpret chromatograms
- deduce, from the given melting points and boiling points, the identities of substances and their purity

1.3 Identification of lons and Gases

- describe the use of aqueous sodium hydroxide and aqueous ammonia to identify the following aqueous cations: aluminium, ammonium, calcium, copper(II), iron(III), iron(III), lead(II) and zinc (formulae of complex ions are not required)
- describe tests to identify the following anions: carbonate (by the addition of dilute acid and subsequent use of limewater), chloride (by reaction of an aqueous solution with nitric acid and aqueous silver nitrate), iodide (by reaction of an aqueous solution with nitric acid and aqueous lead(II) nitrate), nitrate (by reduction with aluminium and aqueous sodium hydroxide to ammonia and subsequent use of litmus paper) and sulphate (by reaction of an aqueous solution with nitric acid and aqueous barium nitrate)
- describe tests to identify the following gases: ammonia (using damp red litmus paper), carbon dioxide (using limewater), chlorine (using damp litmus paper), hydrogen (using a burning splint), oxygen (using a glowing splint) and sulphur dioxide (using acidified potassium dichromate(VI))

USEFUL WEBSITES

http://www.learn.co.uk/default.asp?WCI=Unit&WCU	=3757
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- ☐ http://www.yesmag.bc.ca/projects/paper_chroma.html
- □ http://www.chemistrycoach.com/tutorials-9.htm#Chemistry%20Laboratory
- http://chemistry.about.com/library/weekly/aa020603a.htm
- http://www.kids.net.au/encyclopedia/?p=ch/Chemistry_basic_topics





OVERVIEW OF THE TOPIC

EXPERIMENTAL TECHNIQUES ** key areas

1.1	EXPERIMENTAL DESIGN		• Time	
			Temperature	
			Mass	
			Volume	
		В Соцестон ог	Displacement of water	
		Gas	Displacement of air – upward delivery	
			Displacement of air – downward delivery	
1.2	METHOD OF	& SEPARATION	Filtration	
	PURIFICATION	Techniques	Crystallisation	
	AND ANALYSIS		Simple distillation	
			Fractional distillation	
			Paper chromatography	
		& CRITERIA OF	Melting point determination	
		PURITY	Boiling point determination	
		100		
1.3	IDENTIFICATION	TESTS FOR CATIONS	Aluminium	
	OF IONS AND GASES		Ammonium	
	GASES		Calcium	
			Copper(II)	
			• Iron(II)	
			• Iron(III)	
			• Lead(II)	
			• Zinc	
		& TESTS FOR AHIONS	Carbonate	
			Chloride	
			• lodide	
			Nitrate	
			Sulphate	
		TESTS FOR GASES	Ammonia	
			Carbon dioxide	
			Chlorine	
			Hydrogen	
			Oxygen	
			Sulphur dioxide	

EXPERIMENTAL DESIGN



(a) Time, Temperature and Mass (Fig 1)

Fig 1 shows the apparatus used to measure time, temperature and mass in the laboratory.

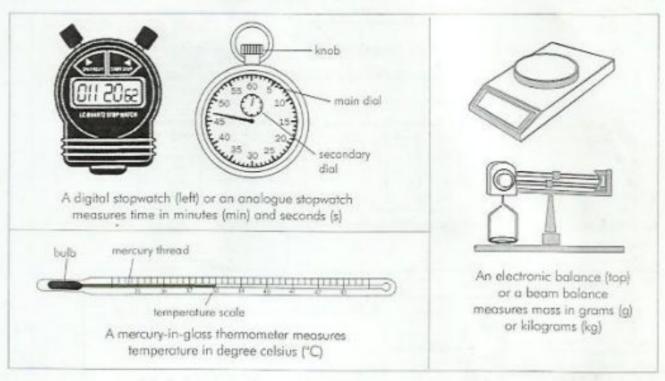


Fig 1 Apparatus for measuring time, temperature and mass

(b) Volumes of Liquids and Gases (Fig 2)

Apparatus for measuring the volumes of liquids vary in accuracy.

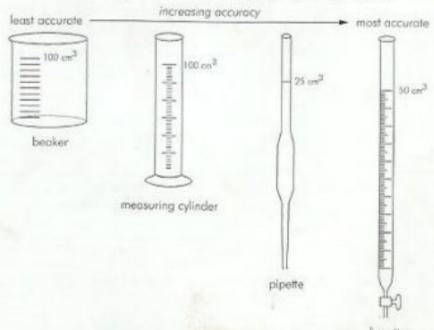


Fig 2 Apparatus for measuring the volumes of liquids

- The beaker is used only for estimating volumes of liquids. If very accurate volumes are required, the burette, which measures up to 0.1 cm³, is used.
- The pipette measures accurately fixed volumes of liquids, e.g. 10.0 cm³, 25.0 cm³. It is accurate to 1 decimal place.
- A graduated gas syringe is used to measure the volumes of gases. (Fig 3)

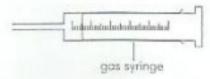


Fig 3 A graduated gas syringe



Always include the units when recording data or carrying out calculations. Missing units will result in the lass of marks.

COLLECTION OF GASES (TABLE 1)

Two factors determine the method used to collect a gas: the density of the gas and the solubility of the gas in water.

Method of Collection	Type of Gases to be Collected
Displacement of water delivery tube gas jar gas	For gases that are insoluble in water, e.g. hydrogen, oxygen.
Displacement of air – upward delivery delivery tube gas ar	For gases that are less dense than air, e.g. hydrogen, ammoria.
Displacement of air – downward delivery gas delivery tube	For gases that are denser than air, e.g. hydrogen chloride, carbon dioxide.

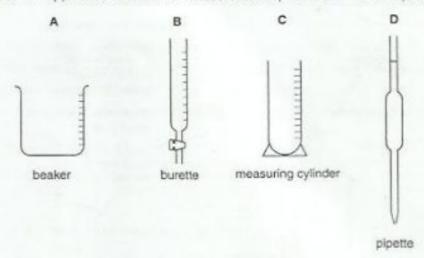
Table 1 Collection of gases



Carbon dioxide gas is sometimes collected by displacement of water. Carbon dioxide is sparingly soluble in water to form carbonic acid. Thus, the valume of gas collected will be less than expected because some carbon dioxide will dissolve into the water.

STOP AND THINK

Which piece of apparatus is used to measure exactly 22.5 cm2 of a liquid?



EXAMTIP 1

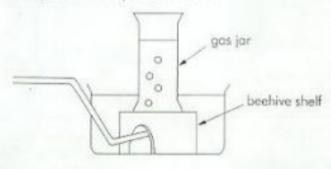
The pipette can also measure volumes of liquids accurately up to one decimal place, but only in fixed volumes, e.g. 10.0 cm³, 25.0 cm³.

[Ans: B] The burette can measure occurately volumes of liquids up to 0.1 cm?

STOP AND THINK

The apparatus shown in the diagram below can be used to collect some gases. Which one of the following gases is most suitably collected by this method?

- A ammonia
- B chlorine
- C hydrogen chloride
- D oxygen
- E sulphur dioxide



EXAMTIP IS

Ammonia will dissolve in water to form an alkaline solution, and should be collected by displacement of air using upward delivery since ammonia is less dense than air. Chlorine, hydrogen chloride and sulphur dioxide will all dissolve in water to form acidic solutions and should be collected by displacement of air using downward delivery since they are all denser than air.

[Ans: D] Oxygen is the only insoluble gas in the list that can be collected by the displacement of water.

METHOD OF PURIFICATION AND ANALYSIS

SEPARADON TECHNIQUES

(a) Filtration (Fig 4)

- Used to separate out an insoluble solid from a liquid, e.g. separating sand from a sand and water mixture.
- Wery fine pores in the filter paper allow the liquid particles to flow through, but retain the large solid particles.

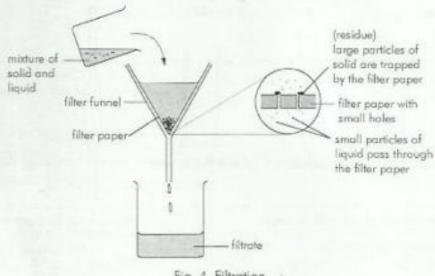


Fig 4 Filtration

(b) Crystallisation (Fig 5)

Used to separate out a pure solid from an impure solution, e.g. separating copper(II) sulphite crystals from impure copper(II) sulphate solution. The impurities will remain dissolved in solution

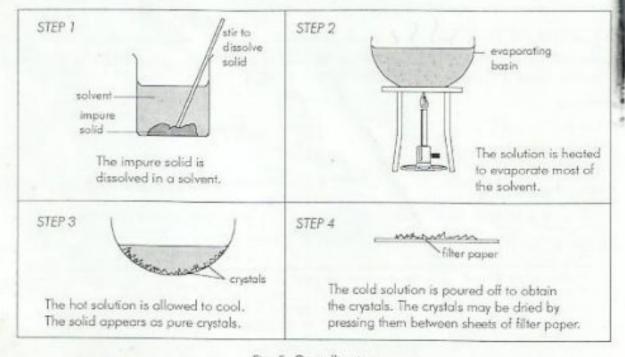


Fig 5 Crystallisation

EXAMTIP IS

Crystallisation must be differentiated from evaporation to dryness. In crystallisation, the solvent is only partially evaporated, leaving a small amount of solution in which the crystals form. Impurities are left behind in the solution when the crystals are filtered off. In evaporation to dryness, all the solvent is removed. The crystals formed may be impure.

STOP AND THINK

Which method is most suitable for obtaining a pure, dry sample of sodium chloride from a mixture of solid sodium chloride and sand?

- A Heat the mixture gently and collect the substance which boils off.
- B Heat the mixture gently and collect the substance that melts.
- C Shake the mixture with water and distil off the liquid.
- D Shake the mixture with water, filter and evaporate the filtrate.

EXAMTIP TOP

If the mixture in the question is solid ammonium chloride and sodium chloride, the method given in answer D cannot be used as both compounds are soluble in water. This mixture can be separated by sublimation. Heat the mixture gently until the solid ammonium chloride starts to sublime, i.e. turn into a gas. Allow the vapour to come into contact with e cold surface so that solid ammonium chloride can reform, leaving solid sodium chloride behind as the residue.

[Ans: D] Shaking with water will dissolve all the sodium chloride. The insoluble sand will be filtered off, leaving sodium chloride solution as the filtrate. Pure, dry sodium chloride can be obtained by evaporation.

(c) Simple Distillation (Fig 6)

- Used to separate a pure liquid from a solution containing dissolved solids, e.g. separating pure water from seawater.
- The pure liquid obtained is called the distillate.

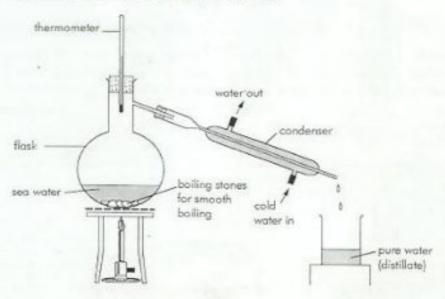
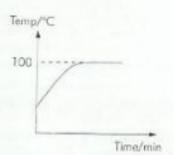


Fig 6 Simple distillation of seawater

- A thermometer is placed at the mouth of the condenser to measure the temperature of the vapour entering it. This temperature is the boiling point of the distillate.
- The graph on the right shows the change in temperature as the seawater is being heated in the flask. The temperature will remain constant at 100 °C throughout the whole time interval when pure water is being collected.



EXAMTIP IS

Cold water enters the condenser from the bottom to ensure that region is the coldest so that all the vapour entering the condenser will turn into a liquid.

(d) Fractional Distillation (Fig 7)

- Used to separate a pure liquid from a solution of two or more liquids, e.g. ethanol from a mixture of ethanol and water.
- Fractional distillation separates according to boiling points. The liquid with the lowest boiling point will be distilled first, followed by the liquid with the next lowest boiling point. As a rough guide, the boiling points of the liquids to be separated should be at least 20°C apart.

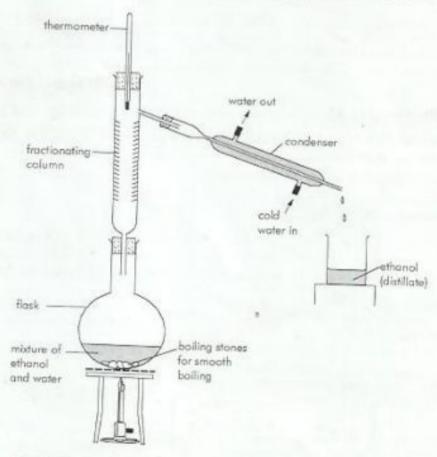
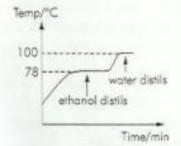


Fig 7 Fractional distillation of a mixture of ethanol and water

The graph on the right shows the change in temperature as the mixture of ethanol and water is being heated in the flask. The temperature will remain at 78 °C when the first distillate (ethanol) is being collected. When all the ethanol has distilled over, the temperature will rise again until it reaches 100 °C. At this temperature, water will be collected as the second distillate.



- Fractional distillation is also used to separate
 - 1 the components of crude oil
 - fermented liquor to obtain alcoholic drinks of a higher concentration (see topic 10)

STOP AND THINK

Which one of the following substances may be separated using a water condenser?

		melting point/°C	boiling point/°C
A	ammonia	-78	-33
В	butane	-135	-0.5
C	sulphur dioxide	-73	-10
D	hydrogen chloride	-115	-85
E	pentane	-130	+36

EXAM TIP IS

The condenser in both the simple and fractional distillation apparatus is used to convert a pure vapour into liquid form. To determine if a gas will condense into a liquid, only the boiling point is considered. The melting points given in the question serve as distractors and is redundant information.

(Ans: E) The temperature of a water condenser is about 25°C. Only pentane has a boiling point above this temperature. Thus, when pentane vapour enters the water condenser, it will condense to form liquid pentane as the distillate.

(e) Paper Chromatography (Fig 8)

Substances in a mixture are separated according to their solubility in the same solvent. The more soluble component in the mixture will tend to remain in the solvent and travel further up the chromatogram, while the less soluble component will separate out onto the paper.

Procedure

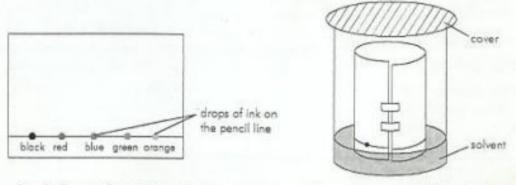


Fig 8 Paper chromatography to separate out the components of black ink

No.1 Experimental Techniques



- 1 Use a pencil to draw the start line.
- Use the black ink sample to make a small dot on the start line, together with some other coloured ink to use as reference.
- ⑤ Fold the paper into a cylinder and place it into a beaker containing the solvent, ensuring that the start line is above the solvent level. Cover the beaker while the chromatogram develops.
- Remove the chromatogram from the beaker just before the solvent reaches the top of the paper.

EXAMTIP TO

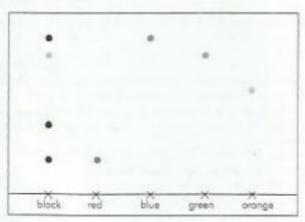
Candidates may be asked to explain why certain steps of an experiment are necessary. In paper chromatography,

(i) the start line must be drawn in pencil. If it is drawn in ink, the components in the ink will also separate out together with the sample dots when the chromatogram is run.

(ii) the start line must be above the solvent level. If it goes below the solvent level, the sample dots will dissolve into the solvent in the beaker instead of travelling up the chromatogram paper.

(iii) the beaker must be covered when the chromatogram is run to reduce evaporation of the solvent from the beaker and to prevent the solvent from evaporating off the paper as it moves up.

Interpretation of results



Dots that has travelled the same distance from the start line in the same solvent belong to the same substance. Hence the black ink contains four colour components – red, blue, green and an unknown colour.

EXAMTIP TO

Paper chromatography has several advantages

(i) It uses only a small amount of sample,

(ii) It can be used to separate heat sensitive substances.

(iii) It can be used to test for the purity of the substance. A pure substance will give only one spot when a chromotogram is run. If it gives two or more spots, then the substance is impure.

STOP AND THINK

It was suspected that a substance X contained one of the three substances P, R or S. Two chromotograms of the four substances were obtained using different solvents. The diagrams show the results.

What does X contain?

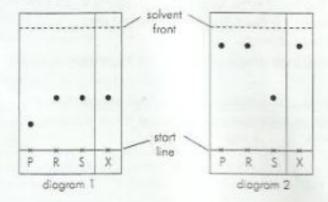
A P only

B R only

C S only

D either P or R

E either R or S



EXAMTIP IN

It is possible for different substances to travel the same distance in the same solvent in paper chromatography, i.e. have the same solubility in the same solvent. To confirm the identity of a substance, another round of chromatography is carried out using a different solvent. If the spots still travel the same distance, then the spots must contain the same substance. We can also conclude from the diagrams that substance X is a pure substance since it only gives one dot in the chromatogram.

[Ans: B] Diagram 1 indicates that X contains either R or S. When the solvent is changed, the chromatogram obtained as shown in diagram 2 indicates that X contains either P or R. Since the same substances will travel the same distance during chromatography even if different solvents are used, X contains only R.

CRITERIA OF PURITY

A pure substance has a fixed melting point and a fixed boiling point.

(a) Determining the Melting Point of a Solid (Fig 9)

- A small sample of the solid is placed into the melting point tube and fied to a thermometer.
- The set-up is placed into a boiling tube containing water and heated slowly. The water is stirred to ensure even heating of the sample.
- Record the temperature when the solid starts to melt
 T_i
- Record the temperature when the solid has fully melted = T_i
- For a pure substance, $T_i = T_i$
- An impure substance will melt over a range of temperatures (T, to T).

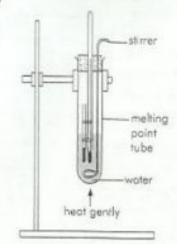


Fig 9 Apparatus for measuring the melting point of a solid

EXAMTED TO

Temperature control is very important when conducting this experiment. If the temperature is increased suddenly, the small sample will melt very fast and T, and T, cannot be determined. Also, if the melting point of the solid is suspected to be above $100^{\circ}C$, a liquid with a boiling point higher than $100^{\circ}C$ will have to be used in place of water.

(b) Determining the Boiling Point of a Liquid (Fig 10)

- The apparatus is set up as shown.
- Anti-bumping granules are added to ensure smooth boiling.
- If the liquid to be tested is flammable, a water bath should be used instead of direct heating with a Bunsen burner.
- Record the temperature when the liquid starts to boil
 T.
- S For a pure substance, T, is a fixed temperature. An impure substance will boil over a range of temperatures.

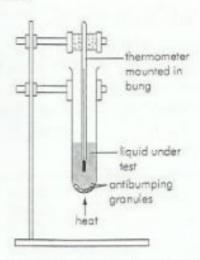


Fig 10 Apparatus for measuring the boiling point of a liquid

EXAMTIP TO

When a substance is impure, its melting point is depressed, i.e. lower than the melting point of the pure substance. Its boiling point will be elevated, i.e. higher than the boiling point of the pure substance. For example, when salt is added to water, its boiling point increases to above 100°C.

STOP AND THINK

A student was given an unknown solid X to identify. He found its melting point to be 133°C. He looked up a chemistry data book and found that a compound called urea had a melting point of 133°C. To find whether X is really urea, he mixed some pure urea and X together. The melting temperature of this mixture was found to be 129 – 132°C. What can be deduced about solid X?

- A X is urea.
- B X is impure urea.
- C X is not urea.
- D X must not be a pure compound.

EXAMTIP IS

Another way to verify if X is urea is to use paper chromatography. Samples of X and pure urea can be dissolved in a suitable solvent and dotted onto a piece of chromatogram paper. The two dats should travel the same distance when a chromatogram is run is X is urea. [Ans. C] If X is urea, then when X is mixed with pure urea and heated, the mixture should melt at 133°C. However, the mixture melted over a range of temperatures, indicating that X is actually an impurity and caused the melting point of pure urea to become depressed. Since X has a fixed melting point, it must be a pure compound, so the only conclusion is that X is not urea.

IDENTIFICATION OF IONS AND GASES

TESTS FOR AQUEOUS CARDINS

- Table 2 summarises the tests for cations.
- When testing for a cation using either aqueous sodium hydraxide or aqueous ammonia, two observations will help identify the cation present:
 - 1 the colour of the precipitate formed on adding a few drops of chemical reagent;
 - 1 the solubility of the precipitate in excess chemical reagent.

Cation	Effect of aqueous sodium hydroxide (NaOH)	Effect of aqueous ammonia (NH ₃)
Aluminium (Al ³⁺)	White precipitate formed. Precipitate is soluble in excess aqueous sodium hydroxide to form a colourless solution.	White precipitate formed. Precipitate is insoluble in excess aqueous ammonia.
Ammonium (NH ₄ +)	No precipitate formed on adding aqueous sodium hydroxide. Ammonia gas produced on warming the mixture.	No chemical reaction.
Caldum (Ca ²⁺)	White precipitate formed. Precipitate is insoluble in excess aqueous sodium hydroxide.	No precipitate formed.
Copper(II) (Cu ²⁺)	Light blue precipitate formed. Precipitate is insoluble in excess aqueous sodium hydroxide.	Light blue precipitate formed. Precipitate is soluble in excess aqueous ammonia to form a dark blue solution.
Iron(II) (Fe ²⁺)	Green precipitate formed. Precipitate is insoluble in excess aqueous sodium hydroxide.	Green precipitate formed, Precipitate is insoluble in excess aqueous ammaria.
Iron(III) (Fe ²⁺)	Reddish brown precipitate formed. Precipitate is insoluble in excess aqueous sodium hydroxide.	Reddish brown precipitate formed. Precipitate is insoluble in excess aqueous ammonia.
Lead(II) (Pb ²⁺)	White precipitate formed. Precipitate is soluble in excess aqueous sodium hydroxide to form a colourless solution.	White precipitate formed. Precipitate is insoluble in excess aqueous ammonia.
Zinc (Zn²+)	White precipitate formed. Precipitate is soluble in excess aqueous sodium hydroxide to form a colourless solution.	White precipitate formed. Precipitate is soluble in excess aqueous ammonia to form a colourless solution.

Table 2 Tests for Cations

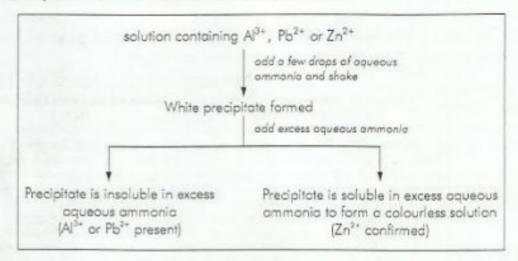
The cations react with the hydroxide ions present in aqueous sodium hydroxide or aqueous ammonia to form insoluble hydroxides. These insoluble hydroxides appear as precipitates.

E.g. Fe^{2+} (aq) + 2OH⁻(aq) \rightarrow $Fe(OH)_2$ (s) from NaOH or NH₂(aq) green precipitate

Some of these precipitates dissolve in excess aqueous sodium hydroxide or aqueous ammonia to form soluble complex solts. These appear as colourless solutions, or in the case of copper(II) ions in excess aqueous ammonia, a dark blue solution.

From table 2,

- Copper(II), iron(II) and iron(III) ions are easily identified by the characteristic colour of their precipitates.
- Aluminium, lead(II) and zinc ions all give the same observations when aqueous sodium hydroxide is used. However, only zinc ions will give a white precipitate soluble in excess aqueous ammonia; aluminium and lead ions do not.



 To distinguish between aluminium and lead(II) ions, dilute hydrochloric acid or aqueous potassium iodide can be used:

$$A^{3+}(aq)$$
 + $3Cl^{-}(aq)$ \rightarrow $AlCl_{3}(aq)$ colourless solution $Pb^{2+}(aq)$ + $2Cl^{-}(aq)$ \rightarrow $PbCl_{2}(s)$ white precipitate

Similar results will be obtained if aqueous potassium iodide is used. Aluminium ions will give a colourless solution of aluminium iodide while lead(II) ions will give a yellow precipitate of lead(II) iodide.

EXAMTIP 5

When there is no visible reaction on adding aqueous sodium hydroxide to an unknown solution, continue the cation test by warming the mixture gently and test for ammonia cas - the cation present may be the ammonium ion.

TESTS FOR AMONS

- Table 3 summarises the tests for anions.
- The tests for the ammonium ion and the nitrate ion produce ammonia gas. Check through the procedure carefully to see if aluminium powder is used the ammonium ion test requires only aqueous sodium hydroxide, but the nitrate test requires both aqueous sodium hydroxide and aluminium powder.

Anion	Test	Test result
Carbonate (CO ₃ ²⁻)	Add dilute acid.	Effervescence. A colourless and odourless gas that forms a white precipitate in limewater is produced. Gas is carbon dioxide.
Chloride (Cl-) [in solution]	Acidify with dilute nitric acid, then add aqueous silver nitrate.	White precipitate of silver chloride formed.
lodide (1-) [in solution]	Acidify with dilute nitric acid, then add aqueous lead(II) nitrate.	Yellow precipitate of lead(II) lodide formed.
Nitrate (NO ₃ -) [in solution]	Add aqueous sodium hydroxide then cluminium foil; warm carefully.	Colourless, pungent gas produced turned moist red litmus paper blue. Gas is ammoria.
Sulphate (SO ₄ 2-) [in solution]	Acidify with dilute nitric acid, then add aqueous barium nitrate.	White precipitate of barium sulphate formed.

Table 3 Tests for Anions

EXAMTIP IS

When recording the observations after conducting tests for the carbonate and the nitrate ion, remember to include the smell and colour of the gas, the chemical test result for the gas as well as the name of the gas. Simply copying from the data sheet provided as 'carbon dioxide produced' or 'ammonia produced' is insufficient and will lead to a loss of marks.

TESTS FOR GASES

- Table 4 summarises the tests for gases.
- Carbon dioxide, sulphur dioxide and chlorine are all acidic gases and will turn moist blue litmus paper red. Hence, the blue litmus paper test is not a conclusive test; it only indicates the presence of an acidic gas. It is necessary to conduct confirmatory tests in order to conclude the presence of a particular gas.
- Ammonia, chlorine and sulphur dioxide have characteristic smells and are thus easily identified.
- When recording observations for gases, it is important to record
 - D presence of effervescence, if any
 - @ colour and smell of the gas
 - 3 chemical test for the gas and test result
 - @ name of the gas

Gas	Colour and Smell of Gas	Test and Test Results
Ammonia (NH ₃)	Colourless, pungent gas	Tums damp red litmus paper blue.
Carbon dioxide (CO ₂)	Colourless, odourless gas	Turns moist blue litmus paper red. Forms white precipitate in limewater.
Chlorine (Cl ₂)	Yellowish green, pungent gas	Bleaches moist red litmus paper. Turns moist blue litmus red, then bleached.
Hydrogen (H ₂)	Colouriess, odouriess gas	Extinguishes a lighted splint with a pop sound.
Oxygen (O ₂)	Colourless, odourless gas	Relights a glowing splint.
Sulphur dioxide (SO ₂)	Colourless, pungent gas	Turns moist blue litmus paper red. Turns orange acidified potassium dichromate(VI) green.

Table 4 Tests for Gases



When testing for hydrogen gas, hold the lighted splint at the mouth of the test tube. When testing for exygen gas, insert the glowing splint into the test tube. When recording the colour change of acidified potassium dichromate(VI), always record as changing from orange to green, and not just acidified potassium dichromate(VI) turned green.

STOP AND THINK

Dilute hydrochloric acid was added to substance X. A gas was produced that turned limewater milky. After the reaction stopped, an excess of aqueous sodium hydroxide was added to the solution and a blue precipitate formed. What was X?

- A CUCO,
- B CuSO
- C FeCO,
- D FeSO,

EXAMTED TOP

If Fe's was present, a green precipitate will be obtained with excess aqueous sodium hydroxide. If $SO_4^{(2)}$ was present, barium nitrate solution will be needed to test for it. No gas would be evalved. Instead, a white precipitate of barium sulphate will be formed. [Ans: A] The gas produced is carbon dioxide, indicating the presence of the $CO_4^{(2)}$ ion. The blue precipitate obtained on adding excess aqueous sodium hydroxide is a test for the $Cu^{(1)}$ ion. The blue precipitate is copper(II) hydroxide.

STOP AND THINK

An aqueous solution containing the compound X was added to a series of solutions and the changes that took place noted.

solution	. observation
iron(III) sulphate	brown precipitate
iron(II) sulphate	green precipitate
zinc sulphate	white precipitate soluble in an excess of the solution of X
magnesium sulphate	white precipitate
copper(II) sulphate	blue precipitate insoluble in an excess of the solution X

What is the compound X?

A aluminium hydroxide

B ammonia

C barium hydroxide

D calcium hydroxide

E sodium hydroxide

EXAMTIP IS

Aluminium hydroxide and benium hydroxide are insoluble in water while calcium hydroxide is only sparingly soluble in water. They will not be able to provide the hydroxide ions needed to precipitate out the insoluble hydroxides of the cations present in the test solutions.

[Ans: E] Notice that all the solutions used to test X are sulphate compounds. Hence the observations obtained must be due to the cations present in the solutions. Cations are usually equeous sodium hydroxide or equeous ammonia. However, X cannot be ammonia because copper(II) ions will give a blue precipitate that is soluble in excess equeous ammonia to form a dark blue solution.

STOP AND THINK

Which one of the following ions will give only a clear solution when heated with an excess of aqueous ammonia?

A A[3+ (eq)

B Zn2+ (aq)

C Fe2+ (aq)

D Fe3+ (og)

E Pb2+ (aq)

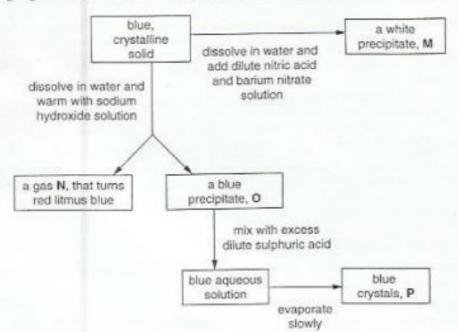
EXAMTED TOP

All and Pbi will form a white precipitate that is insoluble in excess aqueous ammonia, while Fe' and Fe' will form a green and a reddish brown precipitate respectively that is insoluble in excess aqueous ammonia.

[Ans: B] Only Zn² will form a white precipitate that is soluble in excess aqueous ammonia to form a colourless solution

SAMPLE QUESTIONS

- 1. Name a process used to separate a mixture of:
 - (a) three water-soluble dyes
 - (b) water and an insoluble solid
 - (c) two liquids with boiling points 78°C and 100°C
 - (d) water and a dissolved salt
- 2. 25.0 cm³ sodium carbonate solution containing 5.3 g/dm³ was placed in a flask, A few drops of indicator were added. A burette was filled with 0.1 mol/dm³ hydrochloric acid. The acid was titrated against the sodium carbonate solution in the flask. 25.0 cm³ of acid were needed to react with the sodium carbonate. The resulting solution was evaporated until a saturated solution was formed and then allowed to cool. The crystals formed were filtered off, washed and dried.
 - (a) Give the name of the piece of apparatus used to measure 25.0 cm³ of sodium carbonate solution.
 - (b) Suggest how you can tell if a saturated solution was formed.
 - (c) Describe how you would wash and dry the crystals.
 - (d) Give details of a test you could use to show that the crystals contained chloride ions.
- 3. The following figure shows some of the properties of a blue, crystalline solid.



Identify the following.

- (a) substance M
- (b) substance N
- (c) Substance O
- (d) substance P

SOLUTIONS AND EXPLANATIONS

- 1. (a) Paper chromatography
 - (b) Filtration
 - (c) Fractional distillation
 - (d) Crystallisation



Answers should be as precise as possible. For 1(c), the better answer is 'fractional distillation'. We can be referring to either simple distillation or fractional distillation when we just give the answer as 'distillation'. Simple distillation is used to separate out the liquid from a solution containing dissolved solids (e.g. water from salt water), while fractional distillation is used to separate out a liquid from a mixture of miscible liquids (e.g. ethanol from a ethanol-water mixture).

- 2. (a) Pipette
 - (b) A saturated solution is obtained when a thin layer of crystals form as a crust on the surface of the solution.
 - (c) The crystals can be washed with a little cold water and then dried between layers of filter paper.
 - (d) The crystals are first dissolved in distilled water to make a solution. Dilute nitric acid followed by aqueous silver nitrate are then added. If chloride ions are present, a white precipitate of silver chloride will be observed.
- (a) Barium sulphate. Barium nitrate solution is used to test for sulphate ions. The white precipitate formed is barium sulphate.
 - (b) Ammonia. Ammonia is an alkaline gas that will turn moist red litmus paper blue.
 - (c) Copper(II) hydroxide. Sodium hydroxide solution is used to test for cations. The blue precipitate formed is copper(II) hydroxide.
 - (d) Copper(II) sulphate. Copper(II) hydroxide or substance O is a base. It will react with dilute sulphuric acid in a neutralisation reaction to form a salt and water. The salt is copper(II) sulphate, which is obtained by crystallisation.



THE PARTICULATE NATURE OF MATTER

LEARNING OBJECTIVES

Candidates should be able to:

✓ describe the solid, liquid and gaseous states of matter and explain their inter-conversion terms of the kinetic particle theory and the energy changes involved

USEFUL WEBSITES

- □ http://www.chem4kids.com/files/matter_intro.html
- □ http://www.chemistrycoach.com/tutorials-9.htm#Chemistry%20Laboratory
- ☐ http://www.emsb.qc.ca/laurenhill/science/kinetic.html

OVERVIEW OF THE TOPIC

THE PARTICULATE NATURE OF MATTER ** key greas

2.1	THE THREE STATES OF MATTER	Characteristics of the Three States of Matter	
		The Kinetic Theory	
2.2	CHANGES IN STATE	Melting and Freezing	
		Boiling and Condensation	
		Evaporation .	
		Sublimation	

THE THREE STATES OF MATTER

Sours, Liquids and Gases

(a) Characteristics of the Three States of Matter

Matter is anything that occupies space and has mass.

All substances can be classified into 3 states of matter; solid, liquid and gas. Table 1 below shows the characteristics of these 3 states of matter.

	Solids	Liquids	Gases
Shape	Fixed shape	No fixed shape. Takes the shape of the container	No fixed shape. Takes the shape of the container
Volume	Fixed volume	Fixed volume	No fixed valume. Takes the valume of the container
Compressibility	Incompressible	Very slightly compressible, negligible	Very compressible

Table 1 Characteristics of the 3 states of matter

(b) The Kinetic Theory

The Kinetic Theory was proposed to explain the characteristics of the three states of matter. It states that all matter is made up of extremely small particles that are in constant motion. These particles can be atoms, ions or molecules.

	Solid	Liquid	Gas
Diagrammatic representation			00000
Packing between particles	Particles are very dose packed in an orderly arrangement	Particles are close packed in a disorderly arrangement	Particles are far apart in a disorderly arrangement
Forces of attraction between particles	Very strong forces of attraction between particles	Strong forces of attraction between particles	Weak forces of attraction between particles
Movement of particles	Particles can only vibrate and rotate about fixed positions	Particles can move and exchange partners throughout the liquid	Particles are in random motion

Table 2 Differences between the 3 states of matter in terms of the Kinetic Theory

Solids and liquids are incompressible because their particles are all close together. However, gases are compressible because the gas particles are far apart from one another and can be forced to move closer by exerting pressure (Fig. 1).

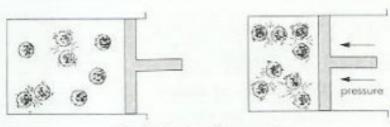


Fig 1 Compressing a gas



When drawing diagrams to represent the particles in a liquid, always draw them close together but in a disorderly manner. Many students tend to draw the particles too far apart as in Fig 3:

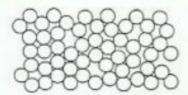


Fig 2 correct

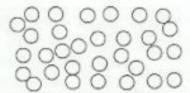


Fig 3 wrong

Remember that liquids are incompressible. The particles in Fig 3 can be further compressed, hence it is more appropriate to use the diagram to represent a gas instead

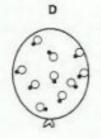
STOP AND THINK

Which diagram shows the arrangement of particles inside a balloon filled with a mixture of helium and argon?









keyhelium atomargon atom



Although argon is denser than helium, B is wrong as the particles of both gases are in continuous motion and will mix to eventually give a homogeneous mixture as shown in A. [Ans: A] Helium and argon are both monoatomic noble gases whose particles are spaced for apart randomly in a container.

STOP AND THINK

An inflated balloon will shrink if it is placed in a refrigerator. What does the lower temperature cause the gas particles in the balloon to do?

- A Move faster and become closer together
- B Move faster and become further apart
- C Move slower and become closer together
- D Move slower and become further apart

EXAMTIP IS

Of the three states of matter, particles in gases have the highest amount of kinetic energy, followed by those in liquids and then solids.

[Ans: C] The gas particles inside the balloon will lose energy to the surroundings on cooling and become less energetic. Hence they will move slower and closer together.

8- CHANGES IN STATE

- Changes in state are physical changes that occur when the particles of a substance absorb or lose energy.
- As a substance is heated, it absorbs energy and changes from a solid to a liquid and finally to a gas. The kinetic energy possessed by its particles increases and they move more vigorously.
- Fig 4 shows the changes in state that can take place.

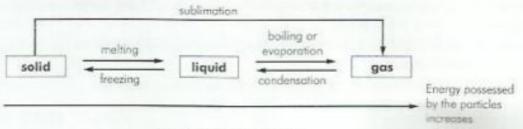


Fig 4 Changes in state

(a) Melting and Freezing

Melting takes place when the particles of a solid absorb enough energy to avercome the forces holding them in fixed positions and move. They rearrange themselves to form a liquid.

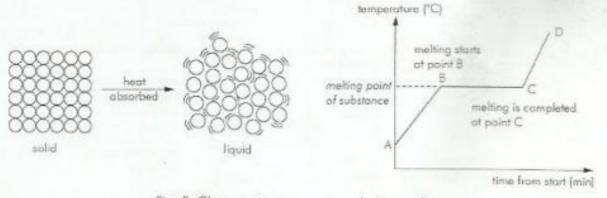


Fig 5 Changes in temperature during melting

- Section AB: The substance remains in solid state. The heat energy provided is absorbed by the solid particles and they vibrate harder about their fixed positions.
- Section BC: A mixture of solid and liquid will be observed at this temperature which is called the melting point of the substance. At point C, the solid has turn completely into a liquid.
- Section CD: The liquid particles continue to absorb heat energy and their kinetic energy increases, causing the temperature of the liquid to rise.

In freezing, the reverse process takes place. At the freezing point, the liquid particles will lose energy when they rearrange to form back the solid structure. This energy loss is equa to the heat removed by cooling, hence the temperature will remain constant.

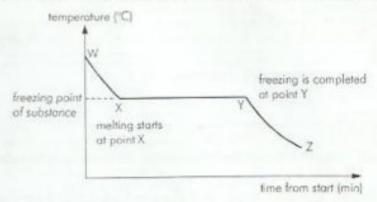


Fig 6 Changes in temperature during freezing.

- Section WX: The substance remains a liquid. As cooling takes place, the particles loses kinetic energy and moves more slowly.
- Section XY: A mixture of liquid and solid is observed as the liquid particles rearrange themselve to form a solid structure.
- Section YZ: The temperature of the solid drops as cooling continues.
- The freezing point is also the melting point for a pure substance.

(b) Boiling and Condensation

Boiling occurs when the particles in a liquid absorb enough energy to overcome the force holding them together and begin to move apart to form a gas.

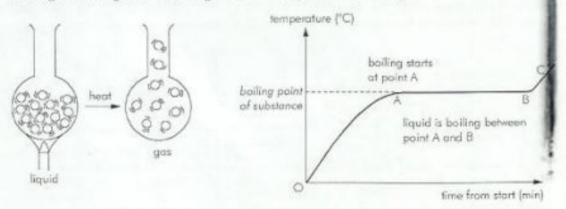


Fig 7 Changes in temperature during boiling

- Section OA: The substance remains in liquid state. The heat energy provided is absorbed by the liquid particles and they move more energetically throughout the liquid.
- Section AB: A mixture of liquid and gas will be observed at this temperature which is calle the boiling point of the substance. At point B, the liquid has turn completely into a gar
- Section BC: The gas particles will absorb energy and move further apart as they become more energetic. The temperature of the gas will rise.
- Condensation is the reverse of boiling. The gas particles will lose energy when cooled an rearrange themselves back into a liquid structure.

(c) Evaporation

- Boiling and evaporation are both physical processes that change a liquid into a gas. The liquid absorbs heat energy during these changes in state.
- Table 3 shows the differences between these 2 processes.

Boiling .	Evaporation		
Occurs at bailing point	Occurs at any temperature below boiling point		
Occurs throughout the liquid	Occurs only at the surface of the liquid		
Bubbles observed No bubbles observed			
Occurs rapidly	Occurs slowly		

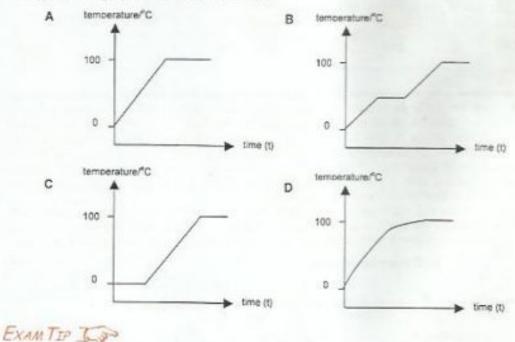
Table 3 Differences between boiling and evaporation

(d) Sublimation

- Sublimation takes place when a solid changes into a gas directly, without going through the liquid state.
- Examples of substances that undergo sublimation include solid carbon dioxide (dry ice), iodine crystals and mothballs.

STOP AND THINK

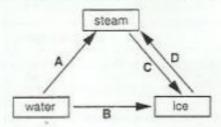
A mixture of water and crushed ice was heated gradually, with constant flame and efficient stirring, until the mixture bailed for a short time. If the temperature of the mixture was plotted against time (t), which graph would be obtained?



Remember that the temperature is constant during any change in state for a pure substance. Answer A will be correct if the experiment started off with water at $0^{\circ}C$. [Ars: C] The experiment started with a mixture of water and crushed ice. Since both solid and liquid states are present, melting must finish before any temperature rise will be observed. Hence, the temperature at the start remains at zero for all the ice to melt before rising.

STOP AND THINK

Which change, A, B, C or D, can involve both condensation and freezing?



[Ans: C] Steam must first be condensed to form water, which then freezes to give ice.

STOP AND THINK

Three substances have the following properties:

substance 1 is brittle;

substance 2 melts at 5°C and boils at 150°C;

substance 3 has a high melting point of 800 °C.

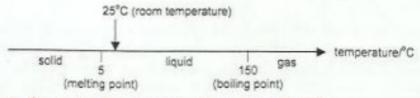
What is the state of each substance at room temperature and pressure?

	substance 1	substance 2	substance 3
Α	gas	liquid	solid
В	solid	gas	solid
C	solid	liquid	solid
D	solid	gas	gas

EXAMTIP TO

Number lines can be used to help determine the state of a substance at a given temperature. Using substance 2 as an example, mark out the melting point and boiling point on the number line and divide it into 3 regions. At temperatures below the melting point, substance 2 is a solid, at temperatures between the melting point and the boiling point, it is a liquid, and at temperatures above the boiling point, it is a gas. To determine the physical state of substance 2 at a given temperature (sey, 25°C), locate this temperature on the number line and see which region it falls into.

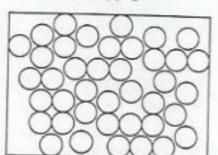
[Ans: C] Substance 1 must be a solid, since the term brittle' means 'easily broken into pieces' and can only be used to describe a solid. At room temperature and pressure (taken as 25° C and 1 atmospheric pressure), substance 2 is a liquid while substance 3 is a solid



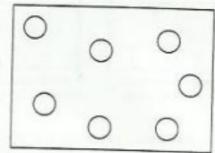
From the number line, substance 2 is a liquid at 25°C,

STOP AND THINK

The diagrams show the arrangement of molecules in a substance at a pressure of 1 atm and at two different temperatures.



at -100°C



Which substance could the diagrams represent?

Substance	Melting Point/°C
A	-183
В	-182
С	-169
D	-114

Boiling Point/°C -89 -162

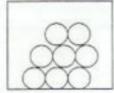
-104 -85

EXAMTIP IS

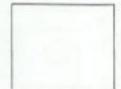
From first impression, it seems that number lines must be drawn to solve the question. This process can be tedious. Always examine the question carefully to see if there is other information provided to solve the problem quickly. MCQs are allocated only 1.5 minutes each and should not require long working steps to solve. [Ans: C] At -150°C, the substance is a liquid. At -100°C, the substance is a gas. This means that the boiling point of the substance is between -150°C and -100°C.

SAMPLE QUESTIONS

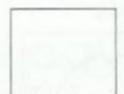
 The drawing in the first box represents the particles in a substance at a temperature below its melting point. Complete the other two boxes.



substance below melting point

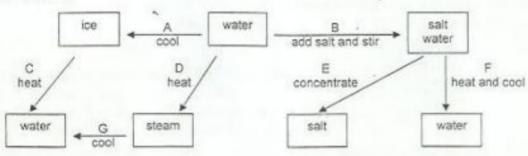


substance just above melting point



substance above boiling point

The following diagram represents changes that can take place to a sample of water. Each change is given a letter.



Write a word to describe the following changes.

- (a) A;
- (d) D;
- (g) G.

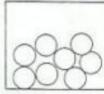
- (b) B;
- (e) E;
- (c) C;
- (f) F

SOLUTIONS AND EXPLANATIONS

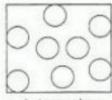
 At temperatures just above the melting point, the substance exists as a liquid. The diagram should show particles close together but in a disorderly arrangement. At temperatures above the boiling point, the substance exists as a gas and the particles will be far apart, occupying all the space in the container.



substance below melting point



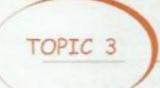
substance just above melting point



substance above boiling point

- 2. (a) A freezing
 - (b) B dissolving
 - (c) C melting
 - (d) D boiling

- (e) E crystallisation
- (f) F simple distillation
- (g) G condensation



ATOMIC STRUCTURE AND CHEMICAL BONDING

LEARNING OBJECTIVES

Candidates should be able to:

3.1 Atomic Structure

- state the relative charges and approximate relative masses of a proton, a neutron and an electron
- describe, with the aid of diagrams, the structure of an atom as containing protons, neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels) — no knowledge of s, p, d and f classification will be expected; a-copy of the Periodic Table will be available in Papers 1, 3 and 5
- define proton number (atomic number) and nuclean number (mass number)
- ☑ interpret and use symbols such as ¹²₆C
- define the term isotopes
- deduce the number of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers

3.2 Structure and Properties of Materials

describe the differences between elements, compounds and mixtures

3.3 Ionic Bonding

- describe the formation of ions by electron loss/gain in order to obtain the electronic configuration of an inert gas
- describe the formation of ionic bonds between metals and non-metals (e.g. NaCl, MgCl₂)
- ☑ relate the physical properties (including electric property) of ionic compounds to their lattice structure

3.4 Covalent Bonding

- describe the formation of a covalent bond by the sharing of a pair of electrons in order to gain the electronic configuration of an inert gas
- describe, using 'dot and cross' diagrams, the formation of covalent bonds between non-metallic elements (e.g. H₂, O₂, H₂O, CH₄ and CO₂)
- ☑ deduce the arrangement of electrons in other covalent molecules
- relate the physical properties (including electrical property) of covalent substances to their structure and bonding

USEFUL WEBSITES

http://www.chemguide.co.uk/atommenu.html#top
http://www.schoolscience.co.uk/content/4/physics/particles/particlesmodel1.html
http://www.krysstal.gom/chemband.html
http://www.beyondbooks.com/psc92/3.asp
http://www.quia.com/dir/chem/
http://www.chem/kide.com/files/atom_structure.html

OVERVIEW OF THE TOPIC

ATOMIC STRUCTURE AND CHEMICAL BONDING Fre key areas

3.1 ATOMIC STRUCTURE		STRUCTURE OF AN	 Relative charges and relative masses of a proton, a neutron and an electron 		
			Proton number (atomic number)		
			Nucleon number (mass number)		
		ARRANGEMENT OF ELECTRONS IN THE ATOM			
3.2 CHEMICAL BONDING		□ Introduction			
	BONDING	& IONEC BONDING	Formation of lons		
			Formation of lanic Bonds		
			Structure of Ionic Compounds		
			Physical Properties of Ionic Compounds		
		& COVALENT BONDING	Formation of Covalent Bonds		
			Structure of Simple Covalent Compounds		
			Physical Properties of Covalent Compounds		
3.3	STRUCTURE AND PROPERTIES OF MATERIALS	MIXTURES AND			

ATOMIC STRUCTURE

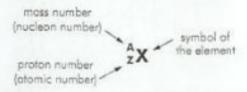
STRUCTURE OF AN ATOM

Atoms are made up of 3 types of particles: protons, neutrons and electrons. The properties of these 3 types of particles are listed in Table 1 below.

Particle	Relative Mass	Relative Charge	Location in Atom
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	0.0005	-1	Electronic shells outside the nucleus

Table 1: Properties of particles found in an atom

- Atoms are electrically neutral. The number of protons is equal to the number of electrons.
- The atomic number (proton number) is defined as the number of protons in an atom. It also gives the number of electrons in the atom.
- The mass number (nucleon number) is defined as the number of protons and neutrons in the atom.
- information about the number of particles in an atom can be found from the Periodic Table:



EXAMTIP TOP

Table 1 is often tested during examinations. When stating the relative charge on the profon or the electron, it is important to write +1 or -1, and not just + or -

STOP AND THINK

Which one of the following atoms has a nucleus which does not contain any neutrons?

A aluminium

D nitrogen

B hydrogen

E oxygen

C lead

EXAMTIR TO

Notice from the above example that the number of neutrons is obtained from the Periodic Table using the top number minus the bottom number. However, this is not true for all cases. For example, chlorine in the Periodic Table is represented as ***CI. This does not mean that the chlorine atom contains 35.5 - 17 = 18.5 neutrons! Chlorine contains 2 isotopes (see later in the chapter), and to account for these 2 isotopes, its mass number is a calculated average value of 35.5.

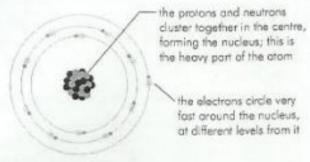
[Ans: B] The number of neutrons in an atom can be easily calculated using the formula number of neutrons = mass number - proton number

	aluminium	hydrogen	lead	nitrogen	oxyger
Information from Periodic Table	27 Al	¹ ₁ H	207 Pb	14N	16 O
mass number	27	1	207	14	16
proton number	13	1	82	7	8
number of neutrons	27-13 =14	1-1 =0	207-82 =125	14-7 -7	16-8 =8

ARRANGEMENT OF ELECTRONS IN THE ATOM

Electrons are arranged in electronic shells (or energy levels) around the nucleus. For the first 20 elements in the Periodic Table,

Shell	Maximum no. of electrons that can occupy the shell	
1 (nearest to the nucleus)	2	
2	8	
3	8	



When drawing electronic structures of atoms, use the symbol of the element to represent the nucleus. The electrons start filling up the shells from the first shell onwards before going on to the next shell.

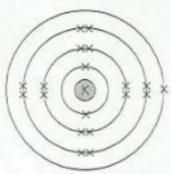
Example 1: lithium, 3Li

The lithium atom contains 3 electrons arranged in 2 shells.



Example 2: Potassium, 39 K

The potassium atom contains 19 electrons arranged in 4 shells.



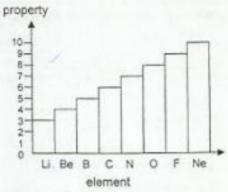
Another way to show the structure of an atom is to write down the number of electrons in each shell. This is known as the electronic configuration of the atom.

Atom	Electronic Configuration
Lithium	2.1
Potasšium	2.8.8.1

Electrons found in the outer shell of an atom are also called valence electrons. These are the electrons that participate in chemical bonding.

STOP AND THINK

The bar chart shows the period of elements from lithium to neon. Which property of these elements is shown on the chart?



- A the number of electrons used in bonding
- B the number of orbits holding electrons
- C the nucleon (mass) number
- D the proton (atomic) number

EXAMTED IS

In the Periodic Table, elements are arranged according to proton number. The horizontal rows are called periods, while the vertical columns are called groups. A new period begins when the maximum number of electrons that can go into an electronic shell is reached. Notice that the period number = number of shells in the element, i.e. for lithium, it is in period 2 and it contains 2 shells, while potassium has 3 shells and is thus located in period 3. One can deduce that for an element located in period 5, for example, iodine, must have 5 shells. Note also that for groups of elements, the group number corresponds to the number of electrons in the outer shell. Hence calcium, found in Group II, has 2 electrons in its outer shell.

[Ans: D] Elements are arranged according to the number of protons in their atoms.

Sotores

- Many elements contain atoms that are slightly different from each other. They have the same proton number but different mass numbers (i.e. same number of protons but different number of neutrons). These atoms are known as isotopes of the element.
- Table 2 below shows the isotopes of some elements.

Element	Name of Isotope	Number of Protons	Number of Neutrons	Number of Electrons
Hydrogen	Protium, H	1	0	1
7.00	Deuterium, ² H	1	1	1
	Tritium, ³ H	1	2	1
Carbon	Carbon-12, 12C	6	6	6
	Corbon-13, 13C	6	7	6
	Carbon-14,14C	6	8	6
Chlorine	Chlorine-35, 35 Cl	17	18	17
	Chlorine-37, 37 Cl	17	20	17

Table 2: Isotopes of some elements

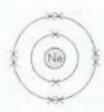
8- CHEMICAL BONDING



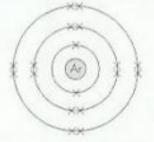
- When atoms combine together in a chemical reaction, we say that a band is formed between the atoms during the reaction. A reactive atom will combine or form bands with other atoms easily, while an unreactive atom will not. The reactivity of an atom depends on its electronic structure.
- Group 0 elements in the Periodic Table are also known as inert gases. They are very stable and hardly form bonds with other elements. The full electronic structure of the first 3 noble gases is shown below:



Helium, He



Neon, 20 Ne



Argon, 18 Ar

- Notice that all the outer electronic shells (or valence shells) are completely filled. This type of arrangement is very stable and makes the atoms inert. Atoms of other elements try to become stable like the noble gases. They share, gain or lose electrons in their valence shells so that they can achieve stable electronic structures.
- There are 2 types of chemical bonds that can be formed between 2 atoms:
 - 1 Ionic bonds valence electrons are transferred from one atom to another
 - O Covalent bands valence electrons are shared

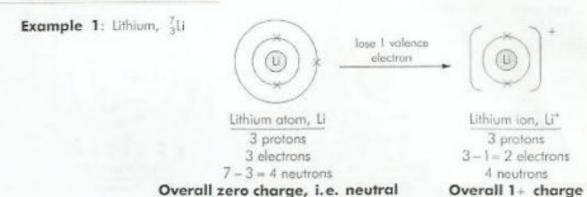


(a) Formation of lons

Atoms can obtain a full outer shell and become stable when they lose or gain valence electrons.

Charged particles called ions are formed.

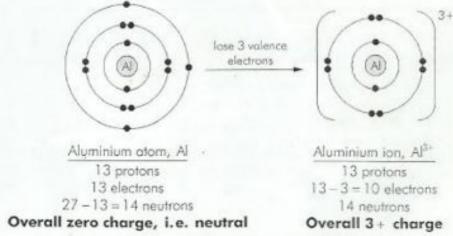
Formation of Positive Ions



The lithium ion now carries a 1+ charge because it has an extra proton. This is represented by enclosing the ion in brackets and writing its charge on the top right hand corner. Note

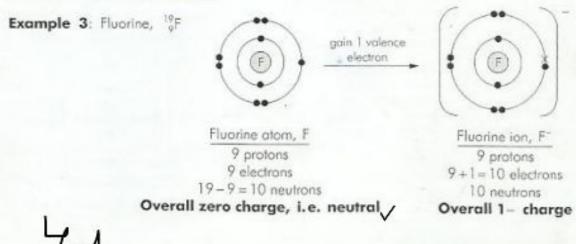
that 1+ is written simply as +. The 1 is not written. We do not draw empty shells either.





In general, when an atom loses n valence electrons to form a stable ion, the ion formed will carry n+ charge.

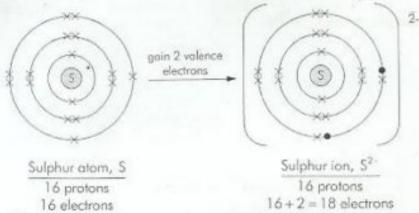
Formation of Negative Ions





- The fluoride ion now carries a 1- charge because it has 1 extra electron. Note that the names of the non-metallic atoms change to end with 'ide' after they become ions. Thus,
 - Fluorine atom becomes fluoride ion
 - Phosphorus atom becomes phosphide ion
 - Oxygen atom becomes oxide ion

Example 4: Sulphur atom, 32S



32-16=16 neutrons Overall zero charge, i.e. neutral 16 neutrons

Overall 2 - charge

In general, when an atom gains m electrons to form a stable ion, the ion will carry mcharge.

STOP AND THINK

Which one of the following ions has the same number of electrons as a neon atom?

	Atomic Number	lon
A	1	hydrogen
В "	11	sodium
C	17	chlorine
D	19	potassium
E	20	calcium

EXAMTIP IN

Elements in the same group (vertical column) of the Periodic Table have the same number of valence electrons. For example, all the elements in Group I have I valence electron each. This implies that elements in Group I will form ions with 1+ charge since they need to lose only 1 valence electron each to form stable structures. In general,

2.1 2.8.1 2.8.8.1

- · Group I elements form 1+ ions
- · Group II elements form 2+ ions
- "Group III elements form 3+ ions

This pattern applies also to negatively charged ions. Hence

- · Group V elements form 3- ions
- · Group VI elements form 2- lens
- · Group VII elements form 1- ions.
- Group IV elements do not usually form ions but reacts through covalent bonding.

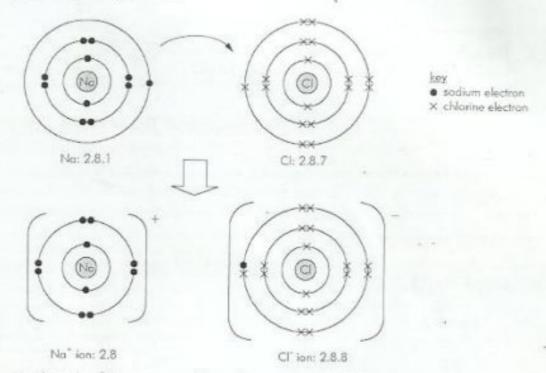
[Ans: B] For atoms, the atomic number (proton number) is the same as the number of electrons. Sodium has the electronic configuration (2.8.1), and so will form No by lesing 1 valence electron and have only 10 electrons, just like a neon atom whose configuration is (2.8).

(b) Formation of Ionic Bonds

- Ionic bonding occurs between metallic and non-metallic atoms. Valence electrons are transferred from the metallic atom to the non-metallic atom so that both atoms achieve a full outer shell and become stable.
- Oppositely charged ions are formed. The metal ion carries positive charge, while the non-metallic ion carries negative charge. These ions attract each other with strong electrostatic forces to form an ionic bond.

Example 1: Sodium Chloride

 The sodium atom transfers 1 valence electron to the chlorine atom so that both atoms achieve a full outer shell.



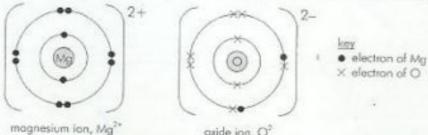
 The formula of the compound formed is written as NaCl. Ionic compounds are electrically neutral, i.e. once the positive and negative ions combine, these charges 'cancel' each other out to give a neutral compound.

EXAMTIP IS

When drawing electrons from different atoms, use different symbols such as * or \times and include a key to describe these symbols. These diagrams are sometimes called 'cross and dot' diagrams. It is also important that the nucleus is clearly labelled with the symbol of the element.

Example 2: Magnesium oxide

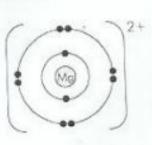
The magnesium atom (2.8.2) transfers 2 valence electrons to the oxygen atom (2.6).



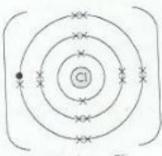
The formula of the compound formed is MgO.

Example 3: Magnesium chloride

The magnesium atom (2.8.2) transfers one valence electron each to 2 chlorine atoms (2.8.7).



magnesium ian, Mg²*



chloride ion, CI



chloride ion, CF

- The compound formed is made up of one Mg2+ ion and 2 CI ions. Hence the formula is written as MgCl₃.
- key
- magnesium electron X chlorine electron

EXAMTIP 100

It is not necessary to draw and show the movement of valence electrons from one atom to another unless the question requires it. Most exam questions will ask for the final structure of the compound only.

STOP AND THINK

The elements X and Y form the compound X,Y. What is the correct electronic configuration of the atoms X and Y?

	electronic co atom of X	infiguration of atom of Y
	2.1	2.7
A B	2.2	2.7
C	2.1	2.6
D	2.2	2.6

[Ans C] The compound contains 2 X' ions and 1 Y' ion. This is worked out from the formula, since 2 ions of X1s needed to 'concel out' the charge on I ion of Y. Since group I elements for ions with +1 charge and group VI ele ments form ions with 2- charge, X mus be from group I with I valence electron, while Y must be from group V with 6 volence electrons.

STOP AND THINK

What is the formula of the compound formed when element X of atomic number 12 combines with element Y of atomic number 17?

A XY.

B X,Y

C XY

EXAMTTE IN

Another way to work out the charges on the ions is to check the groups that X and Y belong to. Since X is in group II, it will form X2, while Y, being in group VII, will form

[Ans: A] Element X (2.8.2) will form X2 while element Y (2.8.7) will form Y. This compound will contain 1 X3 ion and 2 Y ions to be electrically neutral, so the formula is XY

(c) Structure of Ionic Compounds

A solid ionic compound has a giant lattice structure. In this structure, positively and negatively charged ions are held in fixed positions by strong electrostatic forces of attraction.



Fig 1: Arrangement of ions in an ionic lattice

(d) Physical Properties of Ionic Compounds

High melting and boiling points

A lot of energy is required to overcome the strong electrostatic forces of attraction holding the ions in the lattice. Many ionic compounds exist as solids at room temperature.

Conduct electricity when molten or when dissolved in water lonic compounds, when molten or dissolved in water, form mobile ions. These mobile ions are able to move and conduct electricity.

Description Soluble in water, insoluble in organic solvents

STOP AND THINK

Substance X has the following properties. What is X?

- 1. It conducts electricity when molten.
- 2. It has a high melting point.
- 3. It dissolves in an aqueous solution of hydrochloric acid
- A copper
- B ethanol
- C iodine
- D sodium chloride

EXAMTIP IS

Many ionic compounds exist as solids at room temperature because they have very high melting and boiling points due to the strong electrostatic forces of attraction between ions of opposite charge. When heated, the ions in an ionic compound will absorb heat energy, vibrate harder and overcome the forces holding them in fixed positions. In the liquid or molten state, the ions become mobile and can conduct electricity.

[Ans: D] Ethanol is a liquid non-conductor while iodin's sublimes, i.e. there is no liquid or molten state for iodine. Capper does not react or dissolve in aqueous hydrochloric acid, but sodium chloride can dissolve in the water found in aqueous hydrochloric acid.

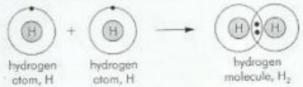
COWNENT BONDERS

(a) Formation of Covalent Bonds

Covalent bonding usually takes place between non-metallic atoms. Valence electrons are shared between these atoms.

Example 1: Hydrogen molecule, H₂

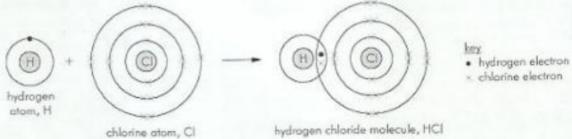
- Electronic configuration of H is 1.
- The hydrogen atom needs 1 more electron in its outermost shell to become stable. This
 electron can be obtained by sharing 1 electron with another hydrogen atom so that both
 atoms have complete outer shells.



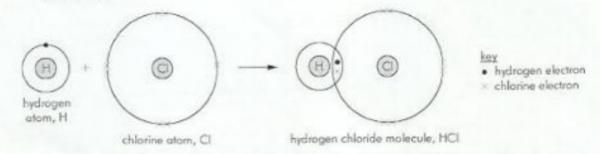
This covalent band can be represented by a dash, so the bonding between the 2 H atoms can be written as H-H. Note that 1 covalent band contains 2 valence electrons, one from each atom.

Example 2: Hydrogen chloride molecule, HCl

- Valence electrons can also be shared between different atoms.
- Electronic configuration of H is 1.
- · Electronic configuration of CI is 2.8.7.

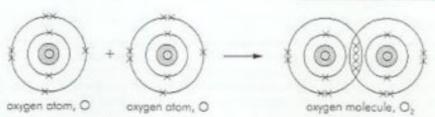


- The covalent bond formed can be represented as H-Cl. When drawing electronic structures
 of different atoms, it is a good idea to represent the electrons from different atoms with
 different symbols (x or •) A key will be needed to interpret these symbols.
- Sometimes, instead of drawing full electronic structures, only outer electronic structures are drawn:



Example 3: Oxygen molecule, O2

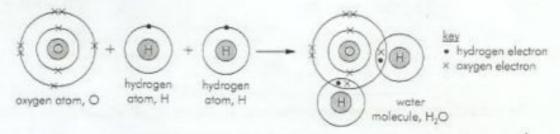
- Electronic configuration of O is 2.6.
- The oxygen atom needs 2 more electrons in its outermost shell to become stable. These
 2 electrons can be obtained by sharing 2 electrons with another oxygen atom:



 The 2 covalent bonds (or the double covalent bond) formed in the oxygen molecule can be written as O = O.

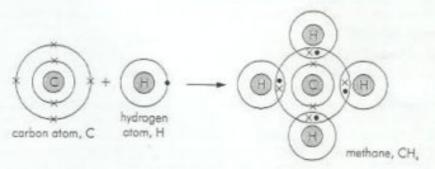
Example 4: Water molecule, H,O

- · Electronic configuration of H is 1.
- Electronic configuration of O is 2.6.
- The oxygen atom will form 2 covalent bonds, one with each hydrogen atom, to obtain a stable structure.



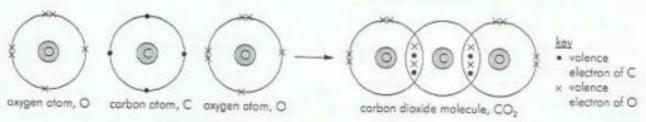
Example 5: Methane, CH,

- Electronic configuration of C is 2.4
- Electronic configuration of H is 1.
- The carbon atom will form 4 covalent bonds, one with each hydrogen atom, to obtain a stable structure.



Example 6: Carbon diaxide, CO₂

- Electronic configuration of C is 2.4.
- Electronic configuration of O is 2.6.
- The carbon atom will form 2 double covalent bonds (i.e. 4 covalent bonds in total), one with each oxygen atom, to obtain a stable structure.



The carbon dioxide molecule can be written as O = C = O.

EXAMTIP TOP

Examine the above examples carefully, and you will notice a pattern - the number of covalent bonds formed by an atom is the number of electrons it needs to obtain a complete outer shell. For example, in its compounds, hydrogen and chlorine will form 1 bond, oxygen will form 2 bonds and carbon will form 4 bonds. This pattern is useful for helping predict the structure of a covalent compound

When drawing covalent structures, always draw the atom that needs to form the most number of bonds in the centre, then add on the rest of the atoms (see Examples 5 and 6)

(b) Structure of Simple Covalent Compounds

Covalent compounds exist as simple molecular structures. The molecules of the compound are held together by weak intermolecular forces that are easily broken by heating.

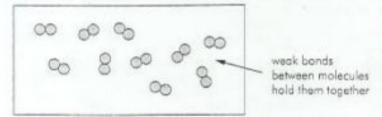


Fig 2: Arrangement of molecules in a simple molecular structure

(c) Physical Properties of Covalent Compounds

Low melting and boiling points

The intermolecular bonds holding the molecules together are weak and easily overcome. Many covalent compounds exist as liquids or gases at room temperature.

Do not conduct electricity in any state

Covalent substances are made up of neutral molecules that are unable to conduct electricity.

Soluble in organic solvents, insoluble in water

E.g. of organic solvents: trichlormethane (CHCl₃), ethanol, petrol

EXAMTED IS

When melting or boiling a covalent compound, only the weak intermolecular forces are broken, not the covalent bonds inside the molecules. Covalent bonds are strong bonds and will require a lot of energy to break. If the covalent bonds in the molecules are broken, then decomposition has taken place and not melting.

STOP AND THINK

The diagram below represents the structure of a malecule. Which one of the following compounds could have the structure shown?

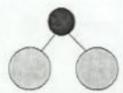
A ammonia

B calcium chloride

C copper(II) oxide

D lead(II) chloride

E water



EXAMTIP TOP

Ionic compounds do not exist as molecules; it is wrong to describe sodium chloride as being made up of molecules of NaCl. NaCl is referred to as the unit of sodium chloride, and it shows the ratio of sedium ions to chloride ions. Hence, a unit of calcium chloride is GoCl, and the ratio of calcium ions to chloride ions is 1:2.

[Ans: E] Since the question states that the structure shown is a molecule, the compound must be covalent, i.e. non-metal atoms combine with non-metal atoms. Ammonia has the formula NH2, while B, C and D are all ionic compounds (metal atoms combine with nonmetallic atoms). Hence the coswer is water, whose formula is H,O.

STOP AND THINK

The table shows the proton (atomic) number of three elements. Which statement about X, Y and Z is correct?

element	proton (atomic) number
X	3
Y	9
Z	10

- A X and Y can combine together to form an ionic compound XY.
- B X and Z can combine together to form an ionic compound XZ.
- C Y and Z can combine together to form a covalent compound YZ.
- D Y is on alkali metal.
- E Z is a halogen

EXAMTIP 13

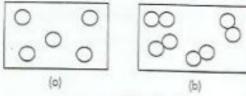
To determine the type of bonding present in a compound, use the general rule - ionic compounds are formed when metallic and non-metallic atoms combine, while covalent compounds are formed when non-metallic atoms combine.

[Ans: A] The electronic configurations of the 3 elements are: X (2.1), Y (2.7) and Z (2.8). Z is a noble gas and will not form compounds. X is a Group I metal and will combine with Y, a Group VII non-metal to form an ionic compound.

STRUCTURE AND PROPERTIES OF MATERIALS

ELEMENTS, MODINES AND COMPOUNDS

Elements are made up of only one kind of atoms. The diagrams in Fig 3 show examples of elements existing as atoms as well as molecules.



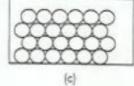
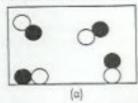
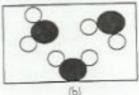


Fig 3: Diagrams showing

- (a) a monoatomic gaseous element made up of atoms, e.g. helium;
- (b) an gaseous element made up of diatomic molecules, e.g. hydrogen gas; and
- (c) a solid element, e.g. iron.

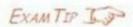
A compound is made up of two or more types of atoms chemically combined together. The diagrams in Fig 4 show examples of some compounds.





- Fig 4: Diagrams showing molecules of a compound made of 2 types of elements, e.g.
 - (a) hydrogen chloride, HCl, and (b) water, H2O
- A mixture is made up of two or more elements or compounds physically combined together. The components can be separated easily from one another using methods such as filtration, a magnet, etc.
- Table 3 summarises the differences between a mixture and a compound.

	Mixture	Compound
Composition	No fixed composition.	Fixed composition. The percentage of each element in a compound is always the same.
Properties	A mixture does not have its own properties but retain those properties of the substances it is made up of. E.g. Saltwater is a mixture of salt and water. It is liquid, like water; and is salty, like pure salt.	A compound has its own set of properties, different from the properties of the elements it is made up of. E.g. Water is a liquid at room temperature, but the elements hydrogen and oxygen are gases.
Melting and Bailing Paints	Melts and boils over a range of temperatures.	Fixed melting and boiling point.
Preparation	No chemical reaction takes place when a mixture is formed.	A chemical reaction takes place when a compound is formed, sometimes with the emission of heat and/or light.
Separation	Easily separated using physical means such as distillation.	Cannot be separated using physical means. Chemical means such as electrolysis are needed.



Elements and compounds are pure substances and will have fixed melting and boiling points. On the other hand, mixtures will melt or boil over a range of temperatures.

STOP AND THINK

Which of the following lists three elements?

- A argon, magnesium, phosphorus
- D magnesium, bronze, hydrogen
- B brass, phospharus, potassium
- petrol, alcohol, water

C chlorine, air, nitrogen

EXAMTIP 500

Brass and bronze are alloys. An alloy is a mixture containing a metal and another element. [Ans: A] Brass and branze are mixtures of elements, air is a mixture of elements and compounds, and petrol is a mixture of compounds. Alcohol and water are pure compounds.

SAMPLE QUESTIONS

1. Use the data sheet to help you answer this question. (a) (i) A calcium ion has an electrical charge. Explain how, and why, this ion becomes charged

when it is formed from an atom.

(ii) Give, using diagrams, the electronic structure of a calcium atom and a calcium ion.

(b) An isotope of calcium has a mass number of 40. Another isotope of calcium has a mass number of 44. What does the nucleus of each isotope contain?

2. Atoms of beryllium and magnesium are represented, respectively, by the symbols:

Be

(a) What do the following numbers tell you about these atoms?

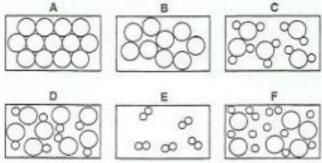
(i) 4 (in the symbol for beryllium)

(ii) 24 (in the symbol for magnesium)

(b) (i) What is the electronic structure of an atom of magnesium?

(ii) In which group of the Periodic Table is magnesium placed?

- (c) How do the electronic structures of beryllium and magnesium indicate that they are both in the same Group of the Periodic Table?
- The drawings represent the particles in six different substances at room temperature and pressure.



Complete the table to show which one of the drawings A to F, best represents each of the following substances. You may use each letter once, more than once or not at all.

substance	copper	a gas	o mixture	hydrogen	water
diagram A - F					

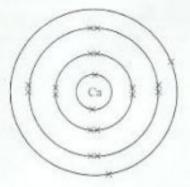
- 4. An element has the electronic structure 2.8.4.
 - (a) Suggest the chemical formula of the compound which this element forms with chlorine. Draw a 'cross and dot' diagram of the structure of this compound.
 - (b) Compare and contrast the boiling point and electrical conductivity of this compound with those of sodium chloride.

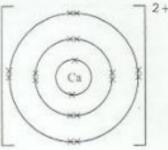
SOLUTIONS AND EXPLANATIONS

1. (a) (i) When the calcium atom undergoes reaction, it loses 2 valence electrons to form the calcium ion. This ion will carry a 2+ charge since the number of protons remain the same; only the number of electrons are affected.

	Atom	lon
No. of protons	20	20
No. of neutrons	20	20
No. of electrons	20	18

(ii)





calcium atom

calcium ion

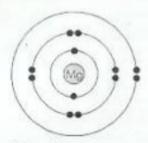
(b) For the isotope with mass number 40:

No. of neutrons
$$= 40 - 20 = 20$$

No, of neutrons =
$$44 - 20 = 24$$

- 2. (a) (i) It is the proton or atomic number of the element. Beryllium has 4 protons.
 - (ii) It is the mass or nucleon number of the element. Magnesium has 24 protons and neutrons.

(b) (i)



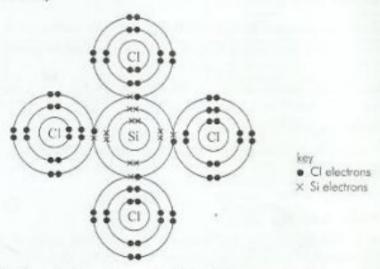
- (ii) Group II
- (c) The structure of beryllium is 2.2, while the structure of magnesium is 2.8.2. Both elements have 2 valence electrons, which indicate that they will be located in Group II of the Periodic Table.

substance	copper	a gas	a mixture	hydrogen	water
diagram A – F	A	E	F	E	C

EXAMTER IS

It is important to be able to identify the substances represented in diagrams such as those given in the question. A is a solid element because the particles are close packed, and there are only one type of atom. B is also an element, but in liquid form. C and D are pure compounds containing two types of atoms. E is a pure gaseous element that exists as diatomic molecules. Finally, F is a mixture of two elements.

(a) The formula is SiCl,.

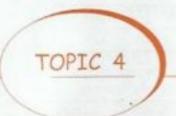


(b) Sodium chloride is an ionic compound with high boiling point. It is an electrical conductor when molten or when dissolved in water.

Silicon tetrachloride is a covalent compound with low bailing point. It does not conduct electricity in all states.

EXAMTIP IS

The silicon atom needs another 4 valence electrons to become stable while the chlorine atoms need 1 valence electron each. Hence the silicon atom will form 4 covalent bonds, while the chlorine atoms will form 1 cavalent bond each. When drawing this structure, place the silicon atom in the centre since it forms the most number of bonds, and then add in the chlorine atoms.



STOICHIOMETRY AND THE MOLE CONCEPT

LEARNING OBJECTIVES

Candidates should be able to:

4.1 Formulae and Equations

- state the symbols of the elements and formulae of the compounds mentioned in the syllabus
- deduce the formulae of simple compounds from the relative numbers of atoms present and vice versa
- deduce the formulae of ionic compounds from the charges on the ions present and vice versa
- ☑ interpret chemical equations with state symbols
- onstruct chemical equations, with state symbols, including ionic equations

4.2 Stoichiometric Calculations

- define relative atomic mass, A
- define relative molecular mass, M
- acalculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24 dm³ at room temperature and pressure); calculations involving the idea of limiting reactants may be set (questions on gas laws and the calculations of gaseous volumes at different temperatures and pressures will not be set)
- ☑ apply the concept of solution concentration (in mol/dm³ or g/dm³) to process the results
 of volumetric experiments and to solve simple problems (simple guidance will be provided
 where unfamiliar reactions are involved)

USEFUL WEBSITES

http://www.creative-chemistry.org.uk/gcse/documents/Module8/N-m08-11.pdf
http://www.marsdenshs.qld.edu.au/subjects/science/junior_science/chem/elements.html
http://www.krysstal.com/reactions.html
http://www.carlton.paschools.pa.sk.ca/chemical/molemass/default.htm
http://www.bbc.co.uk/scotland/education/bitesize/standard/chemistry/calculations/index.shtml
http://www.towson.edu/~ladon/empiric.html
http://www.creative-chemistry.org.uk/gcse/documents/Module7/N-m07-10.pdf
http://www.cchs.co.uk/subjects/science/chemistry/gcse/resources/ionic/ionic.htm
http://www.angelo.edu/faculty/kboudrea/general/formulas_nomenclature/ Formulas_Nomenclature.htm
http://chemed.chem.purdue.edu/genchem/probsolv/stoichiometry/index.html



OVERVIEW OF THE TOPIC

STOICHIOMETRY AND THE MOLE CONCEPT bey areas

4.1	FORMULAE AND EQUATIONS	& CHEMICAL	Chemical formulae of elements
		FORMULAE	Chemical formulae of ionic compounds
			Chemical formulae of covalent compounds
		& CHEMICAL EQUATIONS	Constructing balanced chemical equations with state symbols
			Constructing ionic equations
	- CTOLONG HOLIETTIC	Lancon	
4.2	STOICHIOMETRIC	RELATIVE MASSES	Relative atomic mass
	CALCULATIONS	MASSES	Relative molecular mass
		8 % MASS OF AN ELEMENT IN A COMPOUND	
		THE MOLE CONCEPT	Introduction
			Molar mass
			Determining the number of moles of a substance
			Determining the empirical and molecular formulae of a compound
			Calculating the masses and volumes of gases
			Calculations involving chemical equations
			Calculations involving concentrations of solutions

FORMULAE AND EQUATIONS

CHEMICAL FORMULAS

(a) Chemical Formulae for Elements

- Metals exist as atoms. The chemical formula for a metal is its symbol, e.g. the formula for sodium is Na.
- Most non-metals, with the exception of the noble gases, exist as molecules. Its chemical formula will show both the symbol as well as the number of atoms that make up the molecule. For example, the formula for hydrogen is H₂, where the subscript '2' shows that the molecule is made up of two hydrogen atoms joined together.
- Noble gases exist as atoms. The chemical formula for a noble gas is thus its symbol. For example, the formula for nean is Ne.
- Table 1 below shows the chemical formulae of some common elements.

Metallic Element	Chemical Formula	Non-Metallic Element	Chemical Formula
magnesium	Mg	chlorine	Cl,
zinc	Zn	oxygen	0,
iron	Fe	nitrogen	N,
lead	Pb	carbon	C
manganese	Mn	helium	He
mercury	Hg	argon	Ar

Table 1: Chemical formulae of some common elements

EXAMTIP IN

Sulphur exists as S_a and phosphorus exists as P_a. However, for reasons not explained here, when the formulae of these elements are used in chemical equations, they are simply represented as S and P respectively.

(b) Chemical Formulae of Ionic Compounds

- The formulae of both the positive ion and the negative ion must be determined before the chemical formula of the ionic compound can be written.
- Table 2 below lists some common monoatomic ions with fixed charges.

lon	Symbol	Comments	lon	Symbol	Comments
hydrogen	H ^a		hydride	H-	
lithium	1;+	0	fluoride	F-	
sodium	No+	Group I elements	chloride	CI	Group VII
potassium	K+	form 1 + ions	bromide	Br	elements form
mognesium	Mg ²⁺		iodide	1-	1- ions
calcium	Ca2+	Group II elements	oxide	O2-	Group VI
barium	Bo ²⁺	form 2+ ions	sulphide	S2-	elements form 2- ions
oluminium	All-	Group III elements form 3+ ion	ritride	N ₂	Group V
			phosphide	ps-	3- ions

Table 2: List of common monoatomic ions with fixed charges

Some metals can form positive ions with different charges, depending on the compound that they are found in. Table 3 below lists some common monoatomic ions with variable charges.

Name	Symbol	Name	Symbol
copper(I)	Cu*	mercury(I)	Hg*
copper(II)	Cu ²⁺	mercury(II)	Hg2+
iron(II)	Fe ² *	lead(II)	Pb2+
iron(III)	Fe ¹¹	lead(IV)	Pos
tin(II)	Sn ²⁺	cobalt(II)	Co7+
tin(IV)	Sn**	cobalt(III)	Co3+
chromium(II)	C2:	nickel(II)	N82+
chromium(III)	Cr3+	nickel(IV)	1/2==
manganese(II)	Mn2+	* silver(I)	Ag*
manganese(IV)	Mn ⁴⁺	* zinc(II)	Zn2+

Table 3: List of common monoatomic ions with variable charges

- Note that all these ions are formed from transition metals. The charges that they carry are written in brackets after their names and all the charges are positive.
- Sometimes the charges on silver and zinc ions are not specified. Assume then that the silver ion is Ag*, and the zinc ion is Zn²¹.

EXAMTED IS

It is important that the roman numerals are specified when writing the chemical formula for a compound containing a transition metal. For example, if you are asked to write the formula for iron exide, you can interpret as either iron(II) exide [FeO] or iron(III) exide [FeO]. The two compounds are not the same even though the same elements are found in them. The use of roman numerals helps reduce ambiguity when writing chemical formulae.

Some negative ions exist in groups with an overall charge. Examples of such ions are shown in Table 4. When writing chemical formulae involving such ions, never take apart the ion – take it as a whole group.

Name	Symbol	Name	Symbol
nitrite	NO,	hydroxide	OH-
nitrate	NO.	carbonate	CO.2
sulphite	SO,2	hydrogencarbonate	HCO.
sulphote	SO, 2-	hydrogensulphate	HSO,
manganate(VII)	MnO,	dichromate(VI)	Cr,O,2
phosphate	PO,=	ethanoate	CH,COO
* ammonium	NH,*		

Table 4: List of some common polyatomic ions

Writing chemical formulae for ionic compounds: always remember that an ionic compound is electrically neutral, i.e. total positive charge = total negative charge.

Example 1: sodium chloride

lons present are the sodium ion, Na+ and the chloride ion, Cl-. Since the sodium ion carries 1+ charge and the chloride ion carries a 1- charge, the charges 'cancel' each other out and the compound will contain 1 Na+ and 1 Cl-, i.e. the formula is NaCl.

total 1+ charge
$$\left\{\begin{array}{cc} Na^+ & Cl^- \end{array}\right\}$$
 total 1- charge

Example 2: Potassium oxide

lons present are the potassium ion, K^* and the oxide ion, O^2 . Since the potassium ion carries 1+ charge and the oxide ion carries a 2- charge, 2 potassium ions are needed so that the charges 'cancel' each other out. The compound will contain 2 K^* and 1 O^2 , i.e. the formula is K_*O .

total 2+ charge
$$\left\{ egin{array}{ll} K^+ & O^{2-} \\ K^+ \end{array} \right\}$$
 total 2- charge

Example 3: Magnesium nitride

lons present are the magnesium ion, Mg^{2+} and the nitride ion, N^3 . Since the magnesium ion carries 2+ charge and the nitride ion carries a 3- charge, 3 potassium ions and 2 nitride ions are needed so that the charges 'cancel' each other out. The compound will contain $3 Mg^{2+}$ and $2 N^3-$, i.e. the formula is Mg_*N_* .

total
$$6$$
 + charge $\left\{ \begin{array}{ll} Mg^{2\nu} & N^{3-} \\ Mg^{2+} & N^{3-} \\ Mg^{2+} \end{array} \right\}$ total 6 - charge

Example 4: Calcium sulphate

lons present are the calcium ion, Ca^{2+} and the sulphate ion, SO_4^{-2} . Since the calcium ion carries 2+ charge and the sulphate ion carries 2- charge, the charges will 'cancel' each other out and the compound will contain $1 Ca^{2+}$ and $1 SO_4^{-2}$, i.e. the formula is $CaSO_4$.

total 2+ charge
$$\left\{ \begin{array}{cc} {\sf Ca}^{2^{\frac{1}{4}}} & {\sf SO_4}^2 \end{array} \right\}$$
 total 2- charge

Example 5: Iron(II) nitrate

lons present are the iron(II) ion, Fe^{2*} and the nitrate ion, NO_3^- . Since the iron(II) ion carries 2+ charge and the nitrate ion carries 1- charge, 2 nitrate ions are needed so that the charges will 'cancel' each other out. The compound will contain $1 Fe^{2*}$ and $2 NO_3$, i.e. the formula is $Fe(NO_3)_2$.

total 2+ charge
$$\left\{ \begin{array}{ll} {\rm Fe}^{2+} & {\rm NO_3}^- \\ {\rm NO_3}^- \end{array} \right\}$$
 total 2- charge

EXAMTIP IS

It is not necessary to write brackets around an ion when more than one of it is needed in the formula. In example 3, the formula is written as Mg_N_, not (Mg)_(N)_. However, when more than one anion group is needed, as in example 5, brackets must be used to indicate how many of such groups are present. Hence the formula is Fe(NO₁), and not FeNO₂.

Another technique commonly taught is to 'cross' the charges on ions to obtain the formula

Mg CT - MgCl.

Beware that this technique, although useful, has no conceptual basis. Remember also when using this technique to reduce the subscripts to the lowest terms. Otherwise the formula obtained will be wrong.

STOP AND THINK

What is the formula of uranium(VI) oxide?

A UO

D UO

B U,O,

E UO.

C U,O,

EXAMTIP TO

If the 'cross charges' method has been used here, a possible coreless mistake will be for the student to leave the answer as U.O. Always remember to check the subscripts after writing the formula out using this method to see if they can be reduced to the lowest terms. In this case, the formula should be reduced to UO.

[Ans: E] The ions present are U^o and O^o. Hence the formula will contain 1 U^o and 3 O^o ions in order for the charges to 'cancel' each other out.

(c) Chemical formulae of Covalent Compounds

- Many exceptions exist to the rules for writing the chemical formulae of covalent compounds, making them difficult to remember. Some general rules:
 - Many gases are made up of digtomic molecules, i.e. H₂, O₂, N₃, etc.
 - Group VII elements also exist as diatomic molecules, i.e. Cl₂, Br₂, F₃, etc.
- Table 5 lists some common covalent compounds in the syllabus and their formulae.

Compound	Formula	Compound	Formula
carbon monoxide	CO	nitric acid	HNO,
carbon dioxide	co,	sulphuric acid	H,SO,
sulphur dioxide	SO ₂	*hydrochloric acid	HCI
sulphur trioxide	SO ₃	methane	CH,
ammonia	NH _s	ozone	0,
*hydrogen chloride	HCI	ethanoic acid	СН,СООН
silicon dioxide	SiO ₂	ethanol	C,H,OH

Table 5: List of common covalent compounds

Note that when HCl is in gaseous form, it is called hydrogen chloride gas; when it is dissolved in water, it forms a solution called hydrochloric acid.

CHEWICAL EQUATIONS

(a) Constructing Balanced Chemical Equations with State Symbols

State symbols are written in an equation to identify the physical state of the reactants and products of a chemical reaction. There are 4 state symbols:

Physical State	State Symbol
solid	(s)
liquid or molten	(1)
gas or vapour	(g)
aqueous (dissolved in water)	(aq)

- Note that the 'liquid' state and the 'aqueous' state is not the same. The 'liquid' state of a substance is pure. For a solid substance, the liquid state is obtained by heating the substance until it melts, whereas the 'aqueous' state of a substance is obtained by dissolving it in water.
- Writing a balanced chemical equation with state symbols:
 - ① The word equation must first be given or worked out: E.g. sodium hydroxide + sulphuric acid → sodium chloride + water
 - ? Rewrite each substance using chemical formulae: NaOH + H₂SO₄ \rightarrow Na₂SO₄ + H₂O
 - 3 A balanced chemical equation has the same number of each type of atoms before and after the reaction. In this case, the number of Na and H atoms are not the same on both sides of the equation.

NaOH +
$$H_2SO_a$$
 \rightarrow Na_2SO_a + H_2O
Left Hand Side Right Hand Side *

A '2' must be placed in front of NaOH as well as in front of H2O in order to balance the equation:

① State symbols can be added last to the chemical equation: $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(1)$

EXAMTIP LA

Beginners often find it difficult to balance chemical equations, especially the more complicated ones. A few rules to bear in mind:

If an equation cannot be balanced, it may be wrong. Either the formula of one or more of the substances involved is/are written wrongly or there may be missing/extra substances in the equation.

Never change the chemical formula of compounds when balancing equations. You can only add numbers in front of the chemical formula. For example, 2NaOH and Na,OH has different meanings. 2NaOH means you have 2 units of NaOH (= 2 Na, 2 O and 2 H), while Na,OH means you have 2 Na, 1 O and 1 H.

STOP AND THINK

The equation shows a chemical reaction between sodium and water. The equation is not balanced. What are the values of x and y?

$$xNa + yH_2O \rightarrow 2NaOH + H_2$$

	x	у
A	1	1
В	1	2
C	2	1
n	2	2

EXAMTED TO

No scare is given if you cannot give a balanced equation, even though the formulae of the compounds in your equation are correct. State symbols are not necessary in your balanced chemical equation unless the question requires it.

[Ans: D] The equation is balanced as the number of sodium, oxygen and hydrogen atoms are the same on both sides of the equation.

(b) Constructing Ionic Equations

lonic equations show only the ions involved in a chemical reaction, lons not taking part in the reaction (spectator ions) are 'cancelled' out in the construction of an ionic equation.

Example: copper(II) carbonate + hydrochloric acid

→ copper(II) chloride + carbon dioxide + water

- © Construct the balanced chemical equation: CuCO₃ + 2HCl → CuCl₂ + CO₂ + H₂O
- ② Split only soluble ionic compounds into its ions. Insoluble ionic compounds, elements and covalent compounds remain unchanged. (See topic 6 for solubility rules of common compound)
 CuCO₃ + 2H⁴ + 2Cl⁻ → Cu²⁺ + 2Cl⁻ + CO₃ + H₂O
- O Cancel out the spectator ions. These are the ions that appear in both the left and right side of the equation.

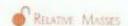
$$C_0CO_3 + 2H^+ + 2\rlap/2l^- \rightarrow C_0^{2+} + 2\rlap/2l^- + CO_2 + H_2O$$

- Rewrite the equation without the spectator ions. This is the ionic equation for the reaction.
 CuCO₁ + 2H⁴ → Cu²⁺ + CO₂ + H₂O
- Compounds that are sparingly soluble or very sparingly soluble can be considered as insoluble when writing ionic equations involving them.

EXAMTED TO

When constructing ionic equations, the number of each particle and the total charge must be the same on both sides of the equation. For example, $Cl_z = 28r^z + Cl_z + 8r_z$ is not a balanced ionic equation since the total charge on the LHS is 2- while the total charge on the RHS is only 1-. The balanced ionic equation will be $Cl_z = 28r^z + 2Cl_z + 8r_z$ where the total charge on both sides of the equation is 2-.

STOICHIOMETRIC CALCULATIONS



(a) Relative Atomic Mass of an Element

The relative atomic mass of an element is the average mass of its atoms expressed on a scale which the carbon-12 atom has a mass of exactly 12.000 units. Figure 1 shows the relative atomic mass scale with the masses of hydrogen, carbon, oxygen and magnesium.

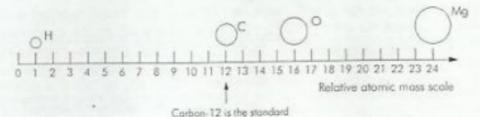


Fig 1: The relative atomic mass scale

- The hydrogen atom is only 1/12 the mass of a carbon-12 atom. Since the carbon-12 atom has a mass of 12 on this scale, the hydrogen atom will have a relative mass of 1. Similarly, the magnesium atom has twice the mass of the carbon-12 atom. It will thus have a relative mass of 24 units.
- All naturally occurring elements are mixtures of isotopes. The relative masses of all the different isotopes present must be taken into account to give a "weighted average" mass for an atom of the element. This is the relative atomic mass of the element.

Example 1: A naturally occurring sample of chlorine contains 75% chlorine-35 and 25% chlorine-37 atoms. Calculate the relative atomic mass of chlorine.

Weighted average mass of a chlorine atom = $(75\% \times 35) + (25\% \times 37) = 35.5$ Thus, the relative atomic mass of chlorine is 35.5, i.e. an average atom of chlorine will have a relative mass of 35.5.

The symbol for relative atomic mass is A_j. The relative atomic mass is a ratio and therefore has no unit.

EXAMTIP 150

There is a clear distinction between mass number and relative atomic mass: the mass number (or nucleon number) of an atom is the number of protons and neutrons in the nucleus of the atom. It is ALWAYS a whole number. The relative atomic mass of an element is the average mass of its atoms compared to the mass of a C-12 atom.

(b) Relative Molecular Mass

The relative molecular mass of an element or a compound is the average mass of its molecules, relative to 1/12 of the mass of a carbon-12 atom.

The symbol for relative molecular mass is M.

M, is calculated by adding together the A, of each atom in the formula. It has no units.

Example 2: Calculate the relative molecular mass of chloroform, $CHCl_3$. M_1 ($CHCl_3$) $\stackrel{*}{\bullet}$ A_1 (C) + A_2 (H) + $3 \times A_2$ (C) = $12 + 1 + 3 \times 35.5 = 119.5$

Example 3: Calculate the relative malecular mass of copper(II) sulphate crystals, CuSO_a.5H₂O.

$$M_{s}$$
 (CuSO₄.5H₂O) = A_s (Cu) + A_s (S) + 9 × A_s (O) + 10 × A_s (H)
= 64 + 32 + 9 × 16 + 10 × 1
= 250

EXAMTIP IS

Strictly speaking, the term 'relative molecular mass' can only be used for substances that exist as molecules, i.e. covalent compounds and some elements. For substances that do not exist as molecules, i.e. ionic compounds, the term 'relative formula mass' is more correct. Fortunately, no distinction is made in the 'O' level examination and the term 'relative molecular mass' is used for all types of compounds.

PERCENTAGE MASS OF AN ELEMENT IN A COMPOUND

The percentage by mass of an element present in a compound is fixed. This percentage can be calculated using the formula

% Mass of element in compound = $\frac{\text{mass of element in compound}}{\text{M, of compound}} \times 100$ $= \frac{\text{No. of atoms} \times \text{A, of element}}{\text{M, of compound}} \times 100$

Example 4: Calculate the percentage by mass of oxygen in carbon diaxide, CO,

% O in
$$CO_2 = \frac{2 \times 16}{12 + 2 \times 16} \times 100\%$$

= 72.7%

EXAMTED 500

If the calculations are correct, the total percentages of all the elements present in a compound should add up to 100%. Hence in example 4, if the % mass of carbon present in carbon diaxide is needed, it is calculated simply as $100 - 72.7 \pm 27.3\%$.

Example 5: Calculate the percentage by mass of water in sodium carbonate crystals, Na_2CO_3 - $10H_2O$

% H₂O in Na₂CO₃.10H₂O =
$$\frac{10 \times (2 \times 1 + 16)}{2 \times 23 + 12 + 3 \times 16 + 10 \times (2 \times 1 + 16)} \times 100\%$$
= 62.9%

The percentage by mass can also be used to calculate the mass of an element in a given sample. The formula is

Mass of element in sample = % of element in compound \times sample mass $= \frac{\text{mass of element in compound}}{M_r \text{ of compound}} \times \text{sample mass}$

Example 6: Calculate the mass of copper in 32 g of copper(II) sulphate.

Mass of Cu =
$$\frac{\text{mass of Cu in CuSO}_4}{\text{M, of CuSO}_4} \times 32 \text{ g}$$

= $\frac{64}{64 + 32 + 4 \times 16} \times 32 \text{ g}$
= 12.8 g

STOP AND THINK

It is now known that one form of carbon has molecules of formula C_{io} . What is the relative molecular mass, M_{io} of these molecules?

A 12

C 360

B 60

D 720

EXAMTIP IS

The formula $C_{\rm so}$ indicates that the molecule contains 60 carbon atoms bonded together. Hence the relative molecular mass of $C_{\rm so}$ is $60 \times A_{\rm s}(C) = 720$. [Ans. D] M, of $C_{\rm so} = 60 \times 12 = 720$

STOP AND THINK

What is the maximum mass of chromium, Cr, that can be extracted from 76 g of chromium(III) exide, Cr, O, 2

A 48 g

C 104 g

B 52 g

D 152 g

EXAMTIP 300

An assumption that the chromium(III) exide sample is 100% pure is made in order to complete the calculation. In practice, it is possible to obtain less than $52\,g$ of chromium if the chromium(III) exide sample is impure.

[Ans: B] Mass of chromium = $\frac{2 \times A_r(Cr)}{M_r(Cr_2O_3)} \times 76 \text{ g} = \frac{2 \times 52}{2 \times 52 + 3 \times 16} \times 76 \text{ g} = 52 \text{ g}$



(a) Introduction

- The unit used in Chemistry to count particles (atoms, molecules, ions, etc.) is the mole.
- One mole = 6×10^{20} particles. The number 6×10^{23} is called the Avagadro Constant. Fig 2 shows how the mole is used in counting particles.

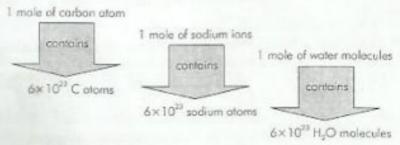


Fig 2: Using the mole to count atoms, malecules and ions

11 The short form for the mole is mol.

(b) Molar Mass

- The molar mass of a substance is the mass of one mol (= 6×10^{23}) of the substance.
- The molar mass of an element made up of atoms is its A_i in grams. For example, the molar mass of aluminium is 27 g, and the molar mass of argon is 40 g.
- The molar mass of a compound or an element made up of molecules is its M, in grams. For example, the molar mass of sodium chloride (NaCl) is 23 + 35.5 = 58.5 g, and the molar mass of hydrogen (H₂) is 2 × 1 = 2 g.

(c) Determining the Number of Moles of a Substance

The number of moles of a substance can be obtained using the formula

Example 7: How many males of atoms are there in 10 g of calcium atoms?

No. of moles =
$$\frac{10}{40}$$
 = 0.25

Example 8: How many males are present in 32 g of sulphur dioxide, SO.2

No. of moles of sulphur dioxide =
$$\frac{32}{32+2\times16}$$
 = 0.5

Use a formula triangle to help remember formulae. If the number of moles is needed, cover the term 'no. of moles' to obtain the equation.



The formula can be manipulated to give the mass of a substance, given the number of moles present. The formula is Mass = Number of moles × Molar mass

This formula can be obtained from the formula by covering the term 'mass' in the formula triangle.

Example 9: What is the mass of 0.1 mol of carbon dioxide molecules? Molar mass of $CO_2 = 12 + 2 \times 16 = 44 \text{ g}$ Mass = $0.1 \times 44 = 4.4 \text{ g}$

Example 10: 0.25 mol of an element X weighs 9.75 g. What is X?

Remember molar mass is the mass of one mol of the substance.

Since 0.25 mol of X weighs 9.75 g, 1 mol of X will weigh 9.75/0.25 = 39 g. From the Periodic Table, X must be potassium.

STOP AND THINK

The formula of copper(I) oxide is Cu₂O. How many grams of oxygen are combined with 64 g of copper in this compound?

A 8 g

C 38 g

B 22 g

D 64 g

EXAMTED 15

The ratio for the above question is obtained from the chemical formula itself, 2 Cu atoms combine with 1 O atom to form a unit of Cu₂O. This means if we have $2 \times (6 \times 10^{12})$ Cu atoms, they will combine with 6×10^{12} O atoms. In other words, since 1 mal = 6×10^{12} , 2 mal of Cu atoms will combine with 1 mal of O atoms.

[Ans: A] No. of moles of Cu = 64/64 = 1. From the formula, 2 mol of Cu atoms will combine with 1 mol of O atoms. Hence 0.5 mol of O atoms will combine with 1 mol of Cu. Mass of O is 0.5 = 16 = 8 g.

STOP AND THINK

What is the ratio of the number of molecules in 71 g of chlorine to the number of molecules in 2 g of hydrogen?

A 71:1

C 2:1

B 71:2

D 1:1

EXAMTED IN

In the above question, it is not necessary to calculate the number of malecules of chlorine and hydrogen 1 mol = 6×10^{23} , so 1 mol of chlorine molecules is 6×10^{23} chlorine molecules, and 1 mol of hydrogen molecules is 6×10^{23} hydrogen molecules. This will give the same ratio of 1 1.

[Ans: D] No. of moles of chlorine molecules, $Cl_1 = 71/71 = 1$. No. of moles of hydrogen molecules, $H_1 = 2/2 = 1$. Hence the ratio is 1:1.

(d) Determining the Empirical and Molecular Formulae of a Compound

- The empirical formula shows the lowest ratio of the different atoms in a compound.
- Remember that the formula of a compound shows the ratio of each type of atom in the compound. For example, if the formula is Na₂O, this means that 2 Na atoms combine with 1 O atom to form the compound. If we count using the mol instead, the formula will show that 2 mol of Na atoms combine with 1 mol of O atom to form the compound.
- Using this idea, the calculation to find the empirical formula involves finding the mal ratio of each type of atom in the compound.

Example 11: What is the empirical formula of a compound with composition of 80% copper and 20% sulphur?

	Cu	S
Mass (g)	80	20
Malar mass (g)	64	32
No. of mol	$\frac{80}{64} = 125$	$\frac{20}{32} = 0.625$
Mol ratio	$\frac{1.25}{0.625} = 2$	$\frac{0.625}{0.625} = 1$

The empirical formula is Cu,S.

Example 12:30 g of silicon oxide contains 14 g of silicon. Find the formula of the compound.

Mass of O in compound =
$$30 - 14 = 16 g$$

	Si	0
Mass (g)	14	16
Molar mass (g)	. 28	16
No. of mol	$\frac{14}{28} = 0.5$	$\frac{16}{16} = 1$
Mol ratio	$\frac{0.5}{0.5} = 1$	$\frac{1}{0.5} = 2$

The empirical formula is SiO2.

EXAMTIP 30

In most cases, the mol ratio obtained in the last step contains whole numbers or nearly whole numbers such as 1.96 which can be easily rounded off to a whole number. However, if the ratio obtained contains numbers such as 2.5, do not round it off but use a multiple instead. For example, if the ratio is 1:25, then the ratio 2:5 should be used in the empirical formula instead of 1:3.

The molecular formula shows the actual number of each kind of atom in a compound. It is a multiple of the empirical formula.

Molecular formula = $(Empirical formula)_n$ where n = 1, 2, 3 ...

(i) To find the value of n, we make use of the equation

Relative molecular mass = $n \times relative$ empirical formula mass

Example 13: The empirical formula of a compound is C₂H₄O. Its relative molecular mass is 88. Find the molecular formula.

Let the molecular formula be (C2H4O),

molecular mass = $n \times empirical$ formula mass

$$88 = n \times (2 \times 12 + 4 \times 1 + 16)$$

$$n = \frac{88}{44} = 2$$

The molecular formula is $(C_2H_4O)_2 = C_4H_8O_2$.

EXAMTIP IS

When calculating the molecular formula, after finding the value of n, do not leave the formula as $(C_2H_4O)_2$ but multiply out the numbers to give the formula C_4H_4O .

- Example 14: A compound of carbon and sulphur has a composition of 15.8% carbon and 84.2% sulphur.
 - (a) Find the empirical formula.
 - (b) The relative molecular mass of the compound is 76. Find the molecular formula.

	C	5
Mass (g)	15.8	84.2
Molar mass (g)	12	32
No. of mol	$\frac{15.8}{12} = 132$	$\frac{84.2}{32} = 2.63$
Mol ratio	$\frac{1.32}{1.32} = 1$	$\frac{2.63}{1.32} = 1.99 \approx 2$

The empirical formula is CS2.

(b) Let the molecular formula be (CS,).

molecular mass = n × empirical formula mass

$$76 = n \times (12 + 2 \times 32)$$

$$n = \frac{76}{76} = 1$$

The molecular formula is $(CS_2)_1 = CS_2$.



It is possible for the value of n to be 1. In this case, the empirical formula is also the molecular formula of the compound.

STOP AND THINK

A compound X contains 50% sulphur and 50% oxygen, by mass. What is the formula of compound X?

A SO

C SO₃

B 50,

D 5,0

EXAMTIP TO

Since the relative molecular mass of compound X was not given, one can assume that the empirical formula is also the molecular formula of compound X and select B as the answer.

[Ans: B] The empirical formula is SO...

5	0
50	50
32	16
50 32=1.5625	50 16=3.125
1,5625 1,5625=1	3.125 1.5625=2
	32 50 32=1.5625 1.5625

(e) Calculating the Masses and Volumes of Gases

One mole of any gas has a volume of 24 dm³ (24 000 cm³) at room temperature (25 °C) and standard atmospheric pressure (1 atm). This volume is called the molar volume of a gas.

The volume of gases is measured using either the dm3 or cm3,

To calculate the number of moles of gas at room temperature and pressure (r.t.p), we use the formula

Number of moles of gas =
$$\frac{\text{volume of gas in cm}^3 \text{ at r.t.p.}}{24\,000 \text{ cm}^3}$$

Number of moles of gas =
$$\frac{\text{volume of gas in cm}^3 \text{ at r.t.p.}}{24 \text{ dm}^3}$$

No. of moles =
$$\frac{0.08 \times 1000}{24000} = 0.003$$

Example 16: In an experiment, when hydrochloric acid was reacted with calcium carbonate at room temperature and pressure, 48 cm³ of carbon dioxide gas was produced. Calculate the number of carbon dioxide molecules evolved.

No. of moles of CO₂ molecules =
$$\frac{48}{24000}$$
 = 0.002
No. of CO₂ molecules = 0.002 × 6 × 10²³ = 1.20 × 10²¹

(L) A formula triangle can also be used to remember this formula:



The formula triangle can be manipulated to calculate the volume of a gas, given the number of males.

Example 17: Calculate the volume of 0.016 g of methane, CH, under room conditions.

No. of mol of CH₄ =
$$\frac{0.016}{12 + 4 \times 1} = 0.001$$

Volume of
$$CH_4 = 0.001 \times 24 = 0.024 \, dm^3$$

Example 18: Calculate the mass of 1.5 dm³ of carbon monoxide, CO, under room conditions.

No. of mol of CO =
$$\frac{1.5}{24}$$
 = 0.0625

Mass of CO =
$$0.0625 \times (12 + 16) = 1.75 g$$

EXAMTIP LO

Gases expand and contract upon heating and cooling. When calculating the volumes of gases, bear in mind that the volume of one mole of gas is $24~\rm{dm^3}$ only at $25^{\circ}C$. If the temperature is lowered to $0^{\circ}C$, the volume of the gas will decrease. At this temperature, one mole of gas will occupy only $22.4~\rm{dm^3}$.

STOP AND THINK

The number of particles in one mole of gas is x. What is the number of chlorine molecules in 35.5 g of gaseous chlorine?

A 0.5x

D 35.5x

Bx

E 71.0x

C 2x

EXAMTIP 3

In this question, x is actually the Avagadra constant, 6×10^{23} . If the volume of chlorine gas is required, then it will be $0.5 \times 24 = 12$ dm³, assuming r.t.p. [Ans: A] No. of moles of chlorine molecules, $Cl_{y} = .35.5/(2 \times 35.5) = 0.5$. Since 1 mol of gas contains x particles, 0.5 mol of chlorine molecules = 0.5×10^{12} molecules.

(f) Calculations Involving Chemical Equations

- Calculations from chemical equations involve five main steps, although not all the steps are required in each set of calculations.
- The steps are:
 - STEP ① Write a balanced chemical equation for the reaction and define the problem to be solved by writing down the information that is given in the question.
 - STEP @ Calculate the no. of moles of the substance whose mass is given.
 - STEP 3 Write down the relevant mole ratio from the chemical equation.
 - STEP @ Calculate the no. of moles of the substance to be found.
 - STEP (3) Calculate the required mass or volume.

Example 19: Calculate the mass of water produced when 4.0 g of methane is completely burnt in oxygen.

The balanced equation	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_{\phi}O$ 4.0 g	
No. of mal of methane	$\frac{\text{mass of methane}}{\text{molar mass of methane}} = \frac{4.0}{12 + 4 \times 1} = 0.25$	
Mol ratio	$\frac{\text{mol CH}_4}{\text{mol H}_2\text{O}} = \frac{1}{2}$	1
No. of mol of water	$\frac{\text{mol CH}_4}{\text{mol H}_2\text{O}} = \frac{1}{2} \Rightarrow \text{mol H}_2\text{O} = \frac{2}{1} \times \text{mol CH}_4$ $\text{mol H}_2\text{O} = \frac{2}{1} \times 0.25 = 0.50$	
Mass of water	no. of moles × molar mass of H ₂ O = 0.50 × (16 + 2 × 1) = 9.0 g	

EXAMTIP IS

The crucial step in these calculations is the determination of the mol ratio from the bolanced equation. In example 19, the balanced equation is $CH_a + 2O_c + CO_c + 2H_cO$. One can interpret the equation as 1 malacule of CH_a reacts with 2 molecules of O_c to produce 1 molecule of CO_c and 2 molecules of H_cO_c if we count using the mol instead (remember the mal is just the number 6×10^{23}), we can interpret the equation as 1 mal of CH_a reacts with 2 mol of O_c to produce 1 mol of CO_c and 2 mol of H_cO_c

Many mol ratios can be written for this equation, depending on what two substances we choose. If we want the mol ratio between CH_4 and O_2 , then mol CH_4 : mol O_3 = 1:2. The mol ratio between O_2 and CO_3 is 2:1. However, only the mol ratio between the substance given and the substance to be found is useful. By defining the problem in step 1, we will know which mol ratio is needed to help solve the problem.

Example 20: Calculate the mass of the solid product obtained when 16.8 g of sodium hydrogencarbonate is heated strongly until there is no further change. The equation for the reaction is

$$2NoHCO_3$$
 (s) $\rightarrow No_2CO_3$ (s) $+ CO_2$ (g) $+ H_2O$ (g)

The balanced equation	$2N_0HCO_3(s) \rightarrow N_{0_2}CO_3(s) + CO_2(g) + H_2O(g)$ 16.8 g ?
No. of mol of NaHCO ₃	$\frac{16.8}{23+1+12+3\times16} = \frac{16.8}{84} = 0.2$
Mol ratio	$\frac{\text{mol Na}_2\text{CO}_3}{\text{mol NaHCO}_3} = \frac{1}{2}$
No. of mol of Na ₂ CO ₃	$\frac{\text{mol Na}_2\text{CO}_3}{\text{mol NaHCO}_3} = \frac{1}{2} \Rightarrow \text{mol Na}_2\text{CO}_3 = \frac{1}{2} \times \text{mol NaHCO}_3$ $= \frac{1}{2} \times 0.2 = 0.1$
Mass of Na ₂ CO ₃	0.1×(2×23+12+3×16)=0.1×106=10.6 g

STOP AND THINK

Limestone decomposes when heated as shown in the equation. Which mass of limestone is needed to produce 84 kg of calcium oxide?

A 150 kg

C 1000 kg

B 840 kg

D 1500 kg

EXAMTIP IS

Be careful of the units used in the question. The molar mass is in grams. Hence, when calculating the number of moles of a substance, the mass of the substance must also be in grams.

The balanced equation	CaCO ₃ → CaO + CO ₂ ? 84 g	
No. of mol of CaO	$\frac{84000}{40+16} = \frac{84000}{56} = 1500$	
Mol ratio	$\frac{\text{mol CaCO}_3}{\text{mol CaO}} = \frac{1}{1}$	
No. of mol of CoCO ₃	$\frac{\text{mol } CaCO_3}{\text{mol } CaO} = \frac{1}{1} \Rightarrow \text{mol } CaCO_3 =$	$\frac{1}{1} \times \text{mol CaO}$ $\frac{1}{1} \times 1500 = 1500$
Mass of CoCO,	1500 × (40+12+3 × 16) = 15000	00 g = 150 kg

STOP AND THINK

Magnesium reacts with hydrochloric acid as shown. What is the volume of hydrogen product at r.t.p. when 0.12 g of magnesium reacts?

Mg + 2HCl → MgCl, + H,

A 56 cm³

C 240 cm³

B 120 cm³

D 480 cm³



The calculation steps in this problem are similar to those in previous examples until t last step where volume instead of mass was required. The formula to use to calculate volume is volume = number of moles * 24 dm³.

[Ans: B]

The balanced equation	Mg + 2HCl → MgCl ₂ + H ₂ 0.12 g ?
No. of mol of Mg	$\frac{0.12}{24} = 0.0050$
Mol ratio	$\frac{\text{mol H}_2}{\text{mol Mg}} = \frac{1}{1}$
No. of mol of Hz	$\frac{\text{mol H}_2}{\text{mol Mg}} = \frac{1}{1} \text{ for mol H}_2 = \frac{1}{1} \times \text{mol Mg}$ $= \frac{1}{1} \times 0.0050 = 0.0050$
Volume of H ₂	0.0050 × 24 = 0.12 dm ³ = 120 cm ³

Sometimes the question requires the student to identify the limiting reactant as part of t solution to the calculation. The limiting reactant is the reactant that, once used up, will cau the reaction to stop.

Example 21: 5.6 g of iron burns in 6.4 g of sulphur to form iron(II) sulphide.

- (a) Calculate and find out which reactant is the limiting reactant
- (b) Calculate the mass of iron(II) sulphide formed after the reaction.
- (c) Calculate the mass of the excess reactant that is left after the reaction.
- (a) No. of mol of Fe = $\frac{5.6}{56}$ = 0.1

No. of mol of S =
$$\frac{6.4}{32}$$
 = 0.2

From equation, 1 mol of Fe reacts with 1 mol of S. This means 0.1 mol of Fe will react with 0.1 mol of S. Hence the limiting reactant is Fe.

(b) 0.1 mol of Fe will produce 0.1 mol of FeS

Mass of FeS =
$$0.1 \times (56 + 32) = 8.8 g$$

(c) S is the excess reactant in this reaction. 0.1 mol of S will react with Fe, leaving (0.2-0.1=0.1) mol of S behind.

Mass of excess
$$S = 0.1 \times 32 = 3.2 g$$

EXAMTIP 33

In part (b) of example 21, the mass of product was calculating using the number of moles of Fe, the limiting reactant. Remember that once the limiting reactant is used up, no more reaction takes place. The amount of product formed will depend on how much limiting reactant was present in the reaction.

STOP AND THINK

The equation for the burning of hydrogen is shown. One mole of hydrogen was burned with one mole of oxygen. What was present after the reaction?

$$2H_{2}(g) + O_{2}(g) \rightarrow 2H_{2}O(g)$$

- A 1 mol of steam only
- B 1 mol of steam + 0.5 mol of oxygen gas
- C 1 mol of steam + 0.5 mol of hydrogen gas
- D 2 mol of steam only

EXAMTIP IS

Another twist to the question will be to ask for the number of moles of gas remaining after the reaction was left to stand for some time. If the reaction mixture was left to cool, the steam formed will condense to give water droplets. The gas left will be 0.5 mol of exygen only.

[Ans: 8] From the equation, 2 mol of hydrogen reacts with 1 mol of oxygen to produce 2 mol of steam. If only 1 mol of hydrogen was burnt, it will react with 0.5 mol of oxygen to produce 1 mol of steam. At the end of the reaction, all the hydrogen will be used up and 0.5 mol of oxygen will be left as excess together with 1 mol of steam as the product.

(g) Calculations involving Concentrations of Solutions

- The concentration of a solution is the amount of a solute dissolved in a unit volume of the solution.
- The concentration of a solution is given in either g/dm³ or mol/dm³.
- To calculate concentration, the following formulae can be used:

Concentration (g / dm³) =
$$\frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3)}$$

or Concentration (mol / dm³) =
$$\frac{\text{no. of moles of solute}}{\text{volume of solution (dm}^3)}$$

EXAMTIP TO

Dissolving 5.0 g of sodium chloride in 100 cm³ of water is not the same as dissolving 5.0 g of sodium chloride in a little water and then topping up to 100 cm³. All matter takes up space, and if 5.0 g of sodium chloride is dissolved in 100 cm³ of water, the volume of the final solution will be slightly more than 100 cm³.

Example 22: A solution of glucose contains 0.45 g of glucose in 75 cm³ of solution. What is the concentration of the glucose solution in g/dm³?

Concentration =
$$\frac{0.45 \text{ g}}{\frac{75}{1000} \text{dm}^3} = 6.0 \text{ g/dm}^3$$

Example 23: What is the concentration in mol/dm² of a solution of sodium hydroxide if it contains 3.5 g of sodium hydroxide in 100 cm³ of solution?

No. of moles of NaOH =
$$\frac{\text{mass}}{\text{molar mass}} = \frac{3.5}{23 + 16 + 1} = \frac{3.5}{40} = 0.0875$$

Concentration =
$$\frac{0.0875 \text{ mol}}{1000 \text{ dm}^3} = 0.875 \text{ mol/dm}^3$$

Sometimes, it is necessary to convert from mol/dm³ to g/dm³. The following formula is used:

Concentration (mol / dm
3
) = $\frac{\text{concentration (g / dm}^3)}{\text{molar mass}}$

Concentration (g/dm3) = concentration (mol/dm3) × molar mass

Example 24: 4.5 g of anhydrous sodium carbonate (Na₂CO₃) was dissolved in distilled water and made up to 500 cm³. Express the concentration of this solution in terms of (a) g/dm³ and (b) mol/dm³.

(a) Concentration =
$$\frac{4.5 \text{ g}}{\frac{500}{1000} \text{ dm}^3}$$
 = 9.0 g/dm³

(b) Molar mass of
$$Na_2CO_3 = 2 \times 23 + 12 + 3 \times 16 = 106 g$$

Concentration in mol/dm³ =
$$\frac{\text{concentration (g / dm}^3)}{106 \text{ g}} = 0.085 \text{ mol/dm}^3$$

EXAMTIP IS

A formula triangle can be used here as well. If the mass of the solute is required, then the formula will be Mass = concentration (in g/dm^2) × volume of solution (in dm^2). Always remember to convert volume of solutions from cm^2 to dm^2 when performing calculations 1000 cm^2 = 1 dm^2



Example 25: The dilute sulphuric acid used in school laboratories usually has a concentration of 2 mol/dm³. What is the mass of H₂SO₄ in 250 cm³ of this acid?

No. of moles = Concentration (mol/dm3) × volume (dm3)

$$=\frac{2 \text{ mol}}{\text{dm}^3} \times \frac{250}{1000} \text{dm}^3 = 0.5$$

Mass of $H_2SO_4 = 0.5 \times (2 \times 1 + 32 + 4 \times 16) = 0.5 \times 98 = 49 g$

STOP AND THINK

Which solution of sodium chloride has the greatest concentration?

- A 0.10 mol NaCl in 0.10 dm3 of solution
- B 0.10 mol NaCl in 1.00 dm3 of solution
- C 0.01 mol NaCl in 100 cm3 of solution
- D 0.01 mol NaCl in 1000 cm3 of solution

[Ans: A] A quick calculation will show that the concentration is 1 mol/dm¹ in A, 0.10 mol/dm² in B and C, 0.01 mol/dm² in D.

SAMPLE QUESTIONS

Fertilisers often contain the compounds with the following formulae:

(NH,),HPO,

Name the compound with the formula KCI.

Calculate the relative molecular mass (M,) of (NH₄)₂HPO₄.

Calculate the percentage by mass of nitrogen in ammonium nitrate, NH,NO3.

(a) Define relative atomic mass.

(b) Tungsten metal, W, is manufactured by reducing tungsten(III) oxide, WO, with carbon.

$$WO_3 + 3C \rightarrow 3CO + W$$

Calculate

- (i) the mass of carbon needed to reduce 116 g of tungsten(III) oxide,
- (ii) the maximum mass of tungsten that can be formed.
- 3. (a) What is meant by molar gas volume and what is its value for carbon monoxide at room temperature and pressure?
 - (b) Hydrogen gas burns in oxygen gas to form water. Calculate
 - (i) the volume
 - (ii) the mass

of oxygen needed to burn 24 dm3 of hydrogen, all volumes being measured at room temperature and pressure.

- When the gas hydrogen sulphide, H,S, is mixed with an excess of oxygen and allowed to react, sulphur dioxide and water vapour are formed.
 - (a) Balance this chemical equation.

$$H_2S + O_2 \rightarrow SO_2 + H_2O$$

- (b) (i) An excess of oxygen is mixed with 100 cm³ of hydrogen sulphide and the mixture is burnt. All volumes are measured at 100 °C. What volume of sulphur dioxide and water vapour are formed?
 - (ii) When the products of the reaction are called to room temperature, the valume decreases by approximately 100 cm³. Explain this decrease in valume.
- Urea, (NH₂)₂CO, is formed by reacting together ammonia NH₃ and carbon dioxide, CO₂.

$$2NH_3(g) + CO_2(g) \rightarrow (NH_2)_2CO(s) + H_2O(l)$$

- (a) (i) How many hydrogen atoms are present in one molecule of urea?
 - (ii) Is the urea represented in the chemical equation a gas, a liquid or a solid?
- (b) What mass of urea can be formed from 68 tonnes of ammonia? Refer to the Periodic Table for relative atomic masses.

A solution of 250 cm³ of ethanol in water contained 23 g of ethanol, C₂H₅OH. Calculate the concentration of the ethanol in g/dm2 and mol/dm2.

SOLUTIONS AND EXPLANATIONS

1. (i) potassium chloride

(ii)
$$M_1 = 2 \times 14 + 9 \times 1 + 31 + 4 \times 16 = 132$$

(iii) % N =
$$\frac{2 \times 14}{2 \times 14 + 4 \times 1 + 3 \times 16} \times 100 = \frac{28}{80} \times 100 = 35$$

 (a) The relative atomic mass is the mass of an atom of an element compared to 1/12 of the mass of an atom of carbon-12.

The balanced equation	$WO_3 + 3C \rightarrow 3CO + W$
No. of mol of WO ₃	$\frac{116}{184 + 3 \times 16} = \frac{116}{232} = 0.5$
Mol ratio	$\frac{\text{mol C}}{\text{mol WO}_3} = \frac{3}{1}$
No. of mol of C	$\frac{\text{mol C}}{\text{mol WO}_3} = \frac{3}{1} \implies \text{mol C} = \frac{3}{1} \times \text{mol WO}_3$ $= \frac{3}{1} \times 0.5 = 15$
Mass of C	15×12=18 a

(ii)
$$\frac{\text{mol W}}{\text{mol WO}_3} = \frac{1}{1} \implies \text{mol W} = \frac{1}{1} \times \text{mol WO}_3$$
$$= \frac{1}{1} \times 0.5 = 0.5$$
Mass of W = 0.5 × 184 = 92 g

 (a) The molar gas volume is the volume occupied by 1 mole of any gas at a given temperature and pressure. The value for carbon dioxide at r.t.p. is 24 dm³.

(b) (i)
$$2H_{2}(g) + O_{2}(g) \rightarrow 2H_{2}O(g)$$

Since all the reactants and products are gases, the mol ratio can be written as the volume ratio.

$$\frac{\text{vol } O_2}{\text{vol } H_2} = \frac{1}{2} \implies \text{vol } O_2 = \frac{1}{2} \times \text{vol } H_2 = 12 \text{ dm}^3$$

Volume of oxygen needed is 12 dm3.

(ii) No. of mol of oxygen =
$$\frac{12}{24}$$
 = 0.5

Moss of oxygen gas = $0.5 \times (2 \times 16) = 16$ g

EXAMTIP 53

The mol ratio can be converted to a volume ratio if the substances involved in the mol ratio are gases.

4. (a)
$$2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$$

(b) (i) Since all the reactants and products are gases, volume ratio is used.

$$\frac{\text{vol SO}_2}{\text{vol H}_2\text{S}} = \frac{2}{2} \implies \text{vol SO}_2 = \frac{2}{2} \times \text{vol H}_2\text{S} = 100 \text{ cm}^3$$

$$\frac{\text{vol H}_2O}{\text{vol H}_2S} = \frac{2}{2} \implies \text{vol H}_2O = \frac{2}{2} \times \text{vol H}_2S = 100 \text{ cm}^3$$

- (ii) When the products are cooled to room temperature, the water vapour will condense. The resulting gas will contain only sulphur dioxide, which has a volume of 100 cm³.
- 5. (a) (i) 4 (ii) solid

The balanced equation	$2NH_3(g) + CO_2(g) \rightarrow (NH_2)_2CO(s) + H_2O(l)$ 68 tonnes ?
No. of mol of NH ₃	$\frac{68000000}{14+3\times1} = \frac{68000000}{17} = 4000000$
Mol ratio	$\frac{\text{mol (NH2)2CO}}{\text{mol NH3}} = \frac{1}{2}$
No. of mol of (NH ₂) ₂ CO	$\frac{\text{mol (NH}_2)_2 \text{CO}}{\text{mol NH}_3} = \frac{1}{2} \Rightarrow \text{mol (NH}_2)_2 \text{CO} = \frac{1}{2} \times \text{mol NH}_3$ $= \frac{1}{2} \times 4000000 = 20000000$
Mass of (NH ₂) ₂ CO	2 000 000 x (2 x 14 + 4 x 1+12+16) = 2 000 000 x 60 = 120 000 000 g = 120 tonnes

EXAMTIP IS

1 tonne = 1000000 g or 1 × 100 g.

6. Concentration of ethanol solution in g/dm³ = $\frac{23 \text{ g}}{250 \text{ dm}^3}$ = 92

Molar mass of ethanol = $2 \times 12 + 6 \times 1 + 16 = 46$ g

Concentration of ethanol solution in mol/dm³ = $\frac{92 \text{ g/dm}^3}{46 \text{ g}} = 2$

EXAMTIP IN

Alternatively, to find the concentration in mol/dm2-

No. of mol of ethanol =
$$\frac{23}{46} = 0.5$$

Concentration of ethanol solution = $\frac{0.5 \text{ mol}}{\frac{250}{1000} \text{dm}^3} = 2$



CHEMICAL REACTIONS

LEARNING OBJECTIVES

Candidates should be able to:

5.1 Speed of Reaction

- describe the effect of concentration, pressure, particle size and temperature on the speeds of reactions and explain these effects in terms of collisions between reacting particles
- ☑ interpret data obtained from experiments concerned with speed of reaction

5.2 Redox Reactions

- ☑ define axidation and reduction (redox) in terms of axygen/hydrogen gain/loss
- define redox in terms of electron transfer and changes in oxidation state
- describe the use of aqueous potassium iodide and acidified potassium dichromate(VI) in testing for oxidising and reducing agents from resulting colour changes

5.3 Energy Changes

- describe the meaning of enthalpy change in terms of exothermic (ΔH negative) and endothermic (ΔH positive) reactions
- describe bond breaking as an endothermic process and bond making as an exothermic process

USEFUL WEBSITES

http://www.chemguide.co.uk/physical/basicratesmenu.html#top
http://www.chemguide.co.uk/inorganic/redoxmenu.html#top
http://www.chemistry.co.nz/redox_begin.htm
http://www.chemistry.co.nz/chemequa.htm
http://www.chem.vt.edu/RVGS/ACT/notes/oxidation_numbers.html
http://www.ausetute.com.au/oxistate.html
http://www.alka-seltzer.com/as/experiment/student_experiment1.htm
http://chemistry.about.com/cs/generalchemistry/a/aa051903a.htm
http://chemistry.about.com/library/weekly/bl091903a.htm
http://www.science-groove.org/SSA/Contest01/chemicalbond.html
http://www.emsb.gc.ca/laurenhill/science/exp.html



OVERVIEW OF THE TOPIC

CHEMICAL REACTIONS ** key areas



5.1	SPEED OF REACTION	& Introduction	
	REACTION	& Measuring the Speed of a Reaction	
		& THE COLLISION THEORY	
		FACTORS AFFECTING THE	Particle size
		Speed of a Reaction	Concentration
			Pressure
			Temperature
5.2	REDOX REACTIONS		
	REACTIONS	Loss/Gain of Oxygen/Hydrogen	
		& Electron Gain/Loss	
		& Changes in Oxidation	Assigning axidation numbers
		NUMBER	Determining oxidation and reduction using oxidation numbers
		∃ Tests For ■ Tes	Defining oxidising and reducing agents
		OXIDISHG/REDUCING AGENTS	Use of aqueous potassium iodide to test for oxidising agents
			Use of acidified potassium dichromate(VI) to test for reducing agents
5.3	ENERGY CHANGES	& INTRODUCTION	
		& EXOTHERMIC REACTIONS	
		& ENDOTHERMIC REACTIONS	
		& ENTHALPY CHANGE	
		BOND BREAKING AND BOND FORMATION	

SPEED OF REACTION

NTRODUCTION

- The speed of a chemical reaction refers to how fast reactants are used up or how fast products are formed in a reaction.
- Different chemical reactions have different speeds.

Examples:

- Reaction of potassium metal with water very fast
- · Rusting of an iron nail in the presence of air and water slow, takes a few days
- Gold reacting with axygen in the air no reaction, speed of reaction of gold with axygen is zero.

MEASURING THE SPEED OF A REACTION

The speed of a reaction is defined as

speed of reaction = change in amount of reactant or product time

We can measure the speed of a reaction by measuring either how quickly a product is obtained or how quickly a reactant is used up.

Example 1: Calcium carbonate (limestone or marble) reacts with dilute hydrochloric acid to produce carbon diaxide gas.

$$CoCO_3$$
 (s) + 2HCl (aq) \rightarrow $CoCl_2$ (aq) + CO_2 (g) + H_2O (l)

We can measure the speed of this reaction by measuring

- the volume of carbon dioxide gas (product) formed over time, or
- the decrease in mass of the system due to the loss of carbon dioxide over time.
- Measuring the speed of a reaction by measuring the volume of gas produced:
 - A graduated gas syringe is used to measure the volume of carbon dioxide gas formed over time (Fig. 1).

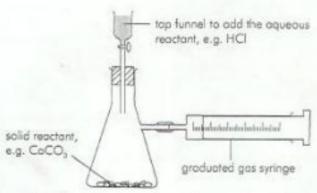


Fig 1: Apparatus set-up for measuring the volume of gas formed in a reaction

The total volume of carbon dioxide gas given off at one-minute intervals is recorded:

Time / min	0	1	2	3	4	5	6
Total valume of CO ₂ / cm ³	0	24	34	38	40	40	40

A graph of total volume of carbon dioxide against time is plotted (Fig 2):

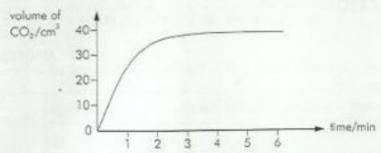


Fig 2: Graph of volume of carbon dioxide against time

1 The gradient at various points of the curve is calculated.

The greater the gradient, the faster the reaction. Fig 3 shows how the speed of the reaction changes over time.

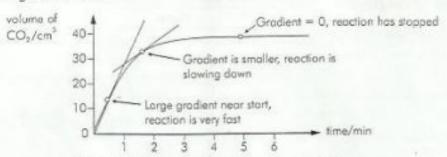


Fig 3: Change in speed of reaction over time

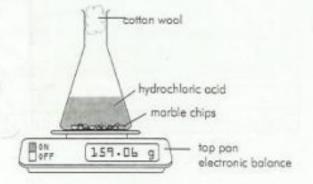
EXAMTIP IN

The reaction is complete once the gradient of the curve becomes zero. In the above example, the reaction is completed in 3.5 minutes.

It is incorrect to say that since the reaction is completed in 3.5 minutes, the reaction is half completed at 3.5/2 = 1.75 minutes. This is because the rate of reaction changes with time - it is faster at the beginning, becomes slower as the reaction proceeds and finally stops. To determine the time when the reaction is half completed, we need to look at how long it takes for half the amount of product to be formed. In this case, it takes approximately one minute for 20 cm² of carbon dioxide gas to form. This time is less than 1.75 minutes.

- Measuring the speed of a reaction by measuring the decrease in mass of system due to loss of product:
 - D A mass balance is used to follow the loss in mass of the system (Fig 4).

Fig 4: Apparatus set-up for measuring loss in mass of system



The mass readings will drop over time as the carbon diaxide gas formed escapes. The mass readings are taken at one-minute intervals and then plotted against time (Fig 5).

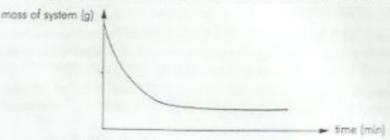
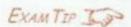


Fig 5: Graph of mass of system against time

The gradient of this graph at various points of the curve will also give the speed of reaction. The reaction is the fastest at the start because the gradient of the graph is the largest. The value of the gradient decreases with time and finally becomes zero. This means that as the reaction proceeds, the reaction slows down and finally comes to a stop.

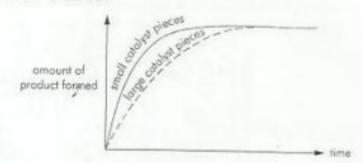


A wad of cotton wool is used as a stopper instead of a rubber bung in Fig 4. The cotton wool will allow the escape of carbon dioxide gas into the atmosphere and prevent the solution inside the conical flask from spraying out. If a rubber bung is used, the pressure inside the flask will build up because carbon dioxide gas cannot escape. There will also be no change in mass reading. Eventually the rubber bung will 'pop' out of the flask when the gas pressure is too great.

STOP AND THINK

A student performs a series of experiments using catalysts. For each experiment he uses the same amounts of reactants and the same amount of catalyst. He plots the graph below from the results. Using only this information, which statement is correct?

- A The catalyst alters both the rate of the reaction and the products formed.
- B The catalyst has no effect on the rate of the reaction.
- C The catalyst produces more product from the same amount of reactant.
- D The reaction is faster when the catalyst pieces are small.



EXAMTIP IN

A catalyst is a substance that can be added to a reaction to make it proceed faster. It does not take part in the reaction and remains chemically unchanged at the end. Therefore, only small amounts of catalyst are needed for a reaction because they can be recycled.

[Ans: D] The gradient of the graph for the experiment using smaller catalyst pieces is steeper than that for the other experiment. Note that both gave the same amount of product. This implies that the reaction is faster using smaller catalyst pieces but the same amount of product will be formed.

THE COUSSION THEORY

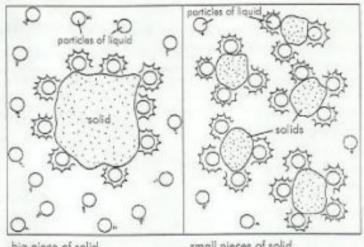
- III For a reaction to occur between 2 particles,
 - 1 the reacting particles must collide with each other, and
 - They must collide with a certain minimum energy known as the activation energy. Otherwise, the particles will just bounce apart after colliding with each other without any change.
- A collision that results in the formation of product molecules is called an effective collision. In general, any factor that increases the number of effective collisions will increase the reaction speed.

FACTORS AFFECTING THE SPEED OF A REACTION

- The speed of a reaction can be altered by 4 factors:
 - ① particle size of reactant smaller particle reactants, faster reactions
 - @ concentration of reactant more concentrated reactants, faster reactions
 - pressure of reactant (for gaseous reactions only) higher pressure, faster reactions
 - 3 temperature higher temperatures, faster reactions

(a) Particle Size

Smaller particles have a greater total surface area compared with larger particles of the same mass (Fig 6). There is more surface area for reaction between the reactants, and the number of effective collisions leading to reaction increases. This increases the speed of reaction.



big piece of solid - small surface area small pieces of solid - large surface area

Fig 6: Comparing the total surface area of a big piece of solid and smaller pieces of the same solid

Example 2: Reaction between (i) marble chips and (ii) powdered calcium carbonate with dilute hydrochloric acid

 $CaCO_{\gamma}(s) + 2HCl(aq) \rightarrow CaCl_{\gamma}(aq) + CO_{\gamma}(q) + H_{\gamma}O(l)$

 If the volume of carbon dioxide gas collected was plotted over time, the following graphs (Fig. 7) will be obtained:

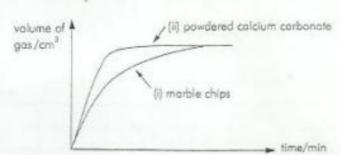


Fig 7: Grophs showing the volume of gas collected against time for the reaction of (i) marble chips and (ii) powdered calcium carbonate with dilute hydrochloric acid

 The graphs show that the reaction using powdered calcium carbonate occurs at a greater speed compared to the reaction using marble chips. The gradient of the curve obtained using powdered calcium carbonate is steeper than that obtained using marble chips for the duration of the reaction.

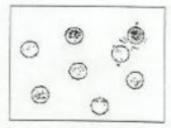
EXAMTIP TO

When comparing reactions, it is important that all other variables are held constant except for the factor that is under study. In this example, the concentration and volume of the acid, the temperature of the surroundings and even the rate of stirring must be held constant while the 2 experiments are conducted so that any difference in results must be due to the difference in particle size of calcium carbonate.

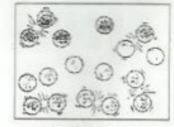
In coal mines and flour mills, workers are not allowed to carry naked lights such as candles or lighted torches because of the fear of explosions. Explosions occur because of the presence of fine particles of flour and coal dust, which burns very easily when mixed with air and ignited.

(b) Concentration

A less concentrated solution has less number of reactant molecules (Fig 8). The frequency of effective collisions decreases, and the speed of reaction goes down.



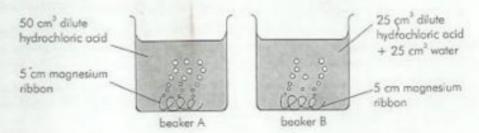
molecules in a less concentrated solution



malecules in a more concentrated solution

Fig 8: Number of molecules in solutions of different concentrations

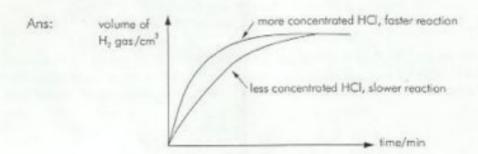
Example 3: Two experiments are conducted to determine the effect of concentration of reactants on the speed of a reaction.



The results of these experiments are shown below.

Beaker	A	В
Time taken for magnesium to dissolve (s)	39	78

- Compare the hydrochloric acid in beakers A and B. Which is the more concentrated acid?
- Ans: The acid in beaker A is twice as concentrated as the acid in beaker B.
- (ii) How would you use the results obtained to explain the relationship between concentration and reaction speed?
- Ans: The reaction in beaker B took twice as long to complete this indicates that reaction speed is decreased when concentration of one of the reactants is decreased.
- (iii) Assuming that all the magnesium is used up, sketch a graph to show how the graph of volume of hydrogen gas collected would vary with time for the reactions in the 2 beakers.



EXAMTIP IN

The hydrochloric acid in beaker A is twice as concentrated in beaker B. If there are x moles of hydrochloric acid in 50 cm³ of it in beaker A, then there is only 0.5x moles of hydrochloric acid in 25 cm³ of it in beaker B.

Concentration of acid in beaker A is
$$\frac{x \text{ moles}}{50 \text{ dm}^3} = 20x \text{ mol}/\text{ dm}^3$$
, while

Concentration of acid in beaker B is
$$\frac{0.5 \times \text{moles}}{1000} = 10 \times \text{mol/dm}^3$$

The volume of hydrogen gas collected in both experiments will be the same because magnesium is the limiting reactant. Since the mass of magnesium ribbon used is the same, the reaction will stop once all the magnesium ribbon is used up.

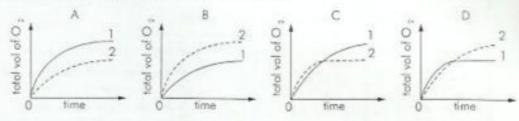
Aqueous hydrogen peroxide decomposes to form water and oxygen.

Two experiments were carried out to measure the rate of production of oxygen from aqueous hydrogen peroxide.

Experiment Solution used

1 100 cm² of 1.0 mol/dm² H₂O₂
2 300 cm² of 0.2 mol/dm² H₂O₀

Which graph best shows the results obtained?



EXAMTER TOP

Aqueous hydrogen peroxide is unstable at room temperature and will decompose slowly on standing. It is also a redox reagent that will turn aqueous acidified potassium dichromate(VI) from orange to green. When aqueous hydrogen peroxide is used during practicals and there is effervescence produced, always test for oxygen gas using a glowing splint. [Ans: A] The graphs have to be analysed in 2 sections - the gradient of the curve, which gives the reaction rate, and the final volume of product obtained. Experiment I will have a faster reaction rate because a higher concentration of hydrogen peroxide is used. Hence the gradient of the curve will be greater. Experiment I will also have more products formed because it contains $0.1~{\rm dm}^3 \times 1~{\rm mol/dm}^4 = 0.1~{\rm mole}$ of hydrogen peroxide, while Experiment 2 contains only $0.3~{\rm dm}^3 \times 0.2~{\rm mol/dm}^3 = 0.06~{\rm mole}$ of hydrogen peroxide.

(c) Pressure

- The higher the pressure, the greater the speed of reaction.
- Changing the pressure affects reactions involving gases only because solids and liquids are incompressible.
- At high pressures, the reactant particles are pushed closer together in a smaller volume (fig 9). This increases the frequency of effective collisions and the speed of reaction increases.

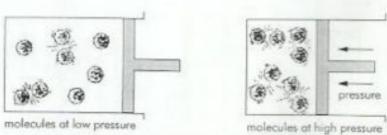


Fig 9: Molecules of gases at different pressures

Example 4: The Haber process used to manufacture ammonia gas uses a high pressure of 200 atm to increase the rate of production of ammonia gas.

 $3H_{2}(g) + N_{2}(g) \rightleftharpoons 2NH_{3}(g)$

(d) Temperature

- lncreasing the temperature increases reaction speed.
- When the temperature is increased, the reactant particles absorb heat energy and become more energetic (Fig 10). The frequency of effective collisions increases and the speed of reaction increases.

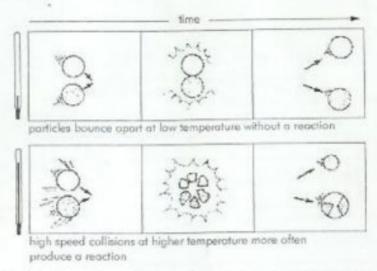


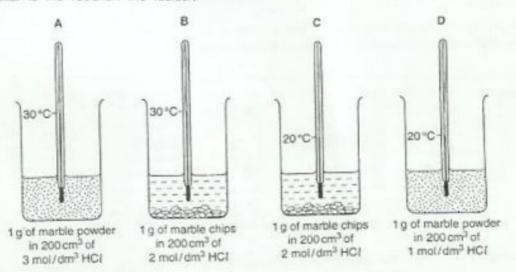
Fig 10: Behaviour of reactant particles at low and high temperatures

Applications

- (i) Food preservation: Low temperatures slows down the chemical reactions that cause food to decay.
- Industrial manufacture: The Contact Process use a high temperature of 450 °C to increase yield of sulphuric acid.

STOP AND THINK

The thermometer shows the temperature at the start of the reaction in each beaker. In which beaker is the reaction the fastest?



EXAMTIP TO

The conditions stated for the experiment in beaker A allow for the most number of effective collisions between the calcium carbonate and hydrochloric acid particles, leading to the fastest reaction rate. The second fastest reaction will occur in beaker B. [Ans: A] Marble powder has greater total surface area, and the concentration and temperature of hydrochlaric acid is also the highest in A.

8 REDOX REACTIONS



- Redox reactions are reactions that involve both oxidation and reduction.
- Oxidation reactions involve
 - 1 gain of oxygen
 - O loss of hydrogen
 - (3) loss of electrons
 - @ gain in oxidation number or oxidation state
- Reduction reactions are the opposite of oxidation reactions. Reduction reactions involve
 - 1 loss of oxygen
 - a gain of hydrogen
 - a gain of electrons
 - loss in oxidation number or oxidation state.
- Oxidation and reduction reactions occur simultaneously. If one reactant is oxidised, then the other reactant must be reduced.

LOSS/GAIN OF OXYGEN/HYDROGEN

- When a substance gains oxygen during a chemical reaction, it is oxidised. If it loses oxygen, the substance is reduced.
 - **Example 1:** $H_2(g) + CuO(s) \rightarrow Cu(s) + H_2O(g)$ H_2 is oxidised to H_2O because it has gained oxygen. CuO is said to be reduced to Cu because it has last oxygen.
 - **Example 2:** $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ CO is oxidised to CO_3 because it has gained axygen. Fe_2O_3 is reduced to Fe because it has lost axygen.
- When a substance loses hydrogen during a chemical reaction, it is oxidised. If it gains hydrogen, substance is reduced.
 - **Example 3:** $H_2S(g) + CI_2(g) \rightarrow 2HCI(g) + S(s)$ H_2S is oxidised to S because it has last hydrogen. CI_2 is reduced to HCI because it has gained hydrogen.
 - **Example 4:** $\text{Cl}_2(g) + \text{N}_2\text{H}_2(g) \rightarrow \text{N}_2(g) + 2\text{HCl}(g)$ N_2H_2 is oxidised to N_2 because it has lost hydrogen. Cl_2 is reduced to HCl because it has gained hydrogen.

The definitions using oxygen and hydrogen are the easiest to use. However, their use is limited to reactions involving oxygen and hydrogen atoms. The most versatile definition by far is the one using oxidation numbers.

ELECTRON GAN/LOSS

- When a substance loses electrons in a chemical reaction, it is oxidised. If it gains electrons, the substance is reduced.
- To use this definition, the ionic equation must be known. Balanced ionic equations involving electrons can be written to show the oxidation and reduction processes. These ionic equations are known as half equations.
 - Example 5: Magnesium reacting with dilute hydrochloric acid

$$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_{+}(g)$$

Mg has lost 2 electrons to form Mg2+. Mg is said to be oxidised to Mg2+.

$$Mg \rightarrow Mg^{2+} + 2e^{-}$$

 $2\mathrm{H}^+$ has gained 2 electrons to form $\mathrm{H_2}$. H^+ is said to be reduced to $\mathrm{H_2}$.

Example 6: Cl, (g) + 2Br (aq)
$$\rightarrow$$
 2Cl (aq) + Br, (l)

2Br has last 2 electrons to form Br., Br is said to be oxidised to Br.,

Cl₂ has gained 2 electrons to form 2Cl. Cl₃ is said to be reduced to Cl.

Note from the examples 5 and 6 that the electrons are simply transferred from one reactant to another. If one substance loses one electron, the other substance must take in the electron during the reaction.

STOP AND THINK

Which statement describes the conversion of a magnesium atom, Mg, to a magnesium ion, Mg2+2

- A The change is reduction, because there is a gain of electrons.
- B The change is reduction, because there is a loss of electrons.
- C The change is oxidation, because there is a gain of electrons.
- D The change is oxidation, because there is a loss of electrons.

EXAMTIP IS

All metals undergo exidation when they lose valence electrons to form stable positive ions. [Ans: D] The magnesium atom loses 2 valence electrons to form the magnesium ion, Mg^{2} . $Mg \rightarrow Mg^{2} + 2e^{-}$

By definition, loss of electrons is oxidation. Mg is oxidised to Mg2.

July 13 power

CHANGES IN OXIDATION NUMBER

(a) Assigning Oxidation Numbers or Oxidation states

- Oxidation numbers are assigned according to the rules below:
 - ① Elements have zero oxidation number.
 - Assume all compounds are ionic and assign the charge the 'ions' carry as their oxidation number.
 - A '+' or a '-' sign must be written in front of the oxidation number, unless it is zero.
- Table 1 below lists the oxidation states of elements in some common substances.

Substance	Oxidation state
hydrogen, oxygen, zinc, iron	O, since all the substances are elements
sodium chloride, NoCl	Na: +1 ; Cl: -1
magnesium fluoride, MgF ₂	Mg: +2 ; F: -1
water, H ₂ O	H: +1; O: -2
carbon dioxide, CO ₂	C: +4 ; O: -2
potassium manganate(VII), KMnO _e	K: +1; O: -2; Mn: +7
potassium dichromate(VI), K ₂ Cr ₂ O ₇	K: +1 ; O: -2 ; Cr: +6
potassium iodide, KI	K: +1 ; I: -1

Table 1: Oxidation states of elements in some common substances

EXAMTIP IN

In determining the oxidation state of manganese in potassium manganate(VII), we have to assume that potassium and oxygen has oxidation numbers +1 and -2 respectively. If we let the oxidation state of manganese be y, we can set up an equation where

(from potassium) (from manganese) (from 4 oxygen) (overall charge of compound)

Solving the equation for y will give us a value of +7. The oxidation state of chromium in potassium dichromate(VI) can be found using the same method. If we let the oxidation state of one chromium be z, then the equation is

 $2 \times (*1)$ *2z $*7 \times (-2)$ = 0 (from 2 potassium) (from 2 chromium) (from 7 oxygen) (overall charge of compound) Solving for z will give a value of *6.

(b) Determining Oxidation and Reduction using Oxidation Numbers

An increase in oxidation number or oxidation state during a reaction is oxidation. A decrease in oxidation number or oxidation state is a reduction.

Example 7: $2Mg(s) + O_{2}(g) \rightarrow 2MgO(s)$

Mg is oxidised as the oxidation state of magnesium has increased from 0 to ± 2 .

 ${\rm O}_2$ is reduced as the oxidation state of oxygen has decreased from 0 to -2.

Example 8:
$$Cr_2O_3 + 2AI \rightarrow AI_2O_3 + 2Cr$$

Cr₂O₄ is reduced as the oxidation state of chromium has decreased from +3

Al is oxidised as the oxidation state of aluminium has increased from 0 to

EXAMTIP IN

Note that in example 8, the definition using exygen can also be used. Cr.O., is reduced to Or because it has lost oxygen, Since exidation and reduction occurs simultaneously, this means that Al must be oxidised to Al,O, by gaining oxygen.

STOP AND THINK

Which underlined substance is reduced?

$$A \quad Ag^+ \ + \ Cl^- \ \rightarrow \ AgCl \qquad C \quad H_2 \ \rightarrow \ 2H^+ \ + \ 2e^-$$

$$C H_2 \rightarrow 2H^+ + 2e$$

$$B \subseteq + CO_2 \rightarrow 2CO$$
 $D \underline{N}_2 + 3H_2 \rightarrow 2NH_3$

EXAM TIP TO

Not all chemical reactions involve exidation and reduction. Reaction A is not a redex reaction because the oxidation numbers of both Ag and Cl have not changed after the reaction. In reaction B, C is axidised because it has gained axygen to form CO, and in reaction C, H₂ is axidised because it has lost electrons to form 2H'.

[Ans: D] In reaction D, the axidation state of nitrogen has decreased from 0 (in N₂) to

-3 (in NH,). N, is reduced to NH,

TESTS FOR ONDISHO/REDUCING AGENTS

(a) Defining Oxidising and Reducing Agents

Oxidising agents are substances that help oxidation take place. In the process, they become reduced. Similarly, reducing agents are substances that help reduction take place. In the process, they become oxidised.

Example 9: $CuS + 4H_2O_2 \rightarrow CuSO_4 + 4H_2O$

CuS is oxidised to CuSO, as it has gained oxygen, and H₂O₂ is reduced to H_oO because it has lost oxygen.

Since H₂O₂ causes CuS to become oxidised (by losing oxygen to it), it is the oxidising agent. On the other hand, CuS is the reducing agent since it causes H₂O₂ to become reduced (by removing oxygen from it).

Example 10:
$$2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl$$

Fe²⁺ is oxidised to Fe³⁺ since the oxidation number of iron has increased from +2 to +3. It is thus the reducing agent for the reaction.

Cl, is reduced to Cl since the oxidation number of chlorine has decreased from 0 to -1. It is thus the axidising agent for the reaction.

The reaction between bromine and iodide ions is shown in the equation. What happens during this reaction?

$$Br_2 + 2l^- \rightarrow 2Br + l_2$$

- A Bromine molecules are the oxidising agent.
- B Bromine molecules are oxidised in bromide ions.
- C lodide ions are reduced to iodine molecules
- D lodide ions gain electrons.

EXAMTIP IS

The oxidation state of iodine has increased from -1 (in Γ) to 0 (in Γ_j), so the iodide ions have been oxidised to iodine molecules and we call the iodide ions the reducing agent for the reaction.

[Ans: A] The exidation number of bromine has decreased from 0 (in Br_j) to -1 (in Br_j), so Br_j molecules are reduced to Br_j ions, Br_j molecules are also the exidising agent for the reaction.

(b) Use of Aqueous Potassium Iodide to Test for Oxidising Agents

Aqueous potassium iodide, KI, is a reducing agent. It contains iodide ions that are easily oxidised to iodine molecules:

When aqueous potassium iodide reacts with an oxidising agent, it will turn from colourless to reddish brown.

EXAMTIP IS

Todine is not very soluble in water. If excess indine is present in solution, it will appear as black precipitate in a reddish brown solution on standing.

(c) Use of Acidified Potassium Dichromate(VI) to Test for Reducing Agents

Acidified potassium dichromate(VI) solution, K₂Cr₂O₂, is an oxidising agent. It contains Cr₂O₂ ions that are easily reduced to Cr²⁺ ions.

$$Cr_3O_3^{2-} \rightarrow Cr^{3+}$$

orange green

When acidified potassium dichromate(VI) solution reacts with reducing agent, it will turn from orange to green.

Suphur dioxide, SO₂, is a reducing agent and its presence is detected by the use of acidified potassium dichramate(VI) solution.

EXAMTIP IN

Acidified potassium manganate(VII), $KMnO_a$, can be used in place of acidified potassium dichromate(VI) as a test for reducing agents. It contains manganate(VII) ions that are reduced to manganese(II) ions during a redex reaction.

The purple manganate(VII) ions are decolourised, i.e. they turn from purple to colourless,

Solution X turns acidified potassium dichromate(VI) from orange to green. What must solution X contain?

A An alkali

C An oxidising agent

B An ammonium salt

D A reducing agent

EXAMTIP IS

If two exidising agents are mixed together, no reaction will take place since exidation and reduction must occur simultaneously. This is also true if two reducing agents are mixed together.

[Ans: D] Solution X must contain a reducing agent.

ENERGY CHANGES



When chemical reactions take place, energy is either taken in or given out from the surroundings in the form of heat and/or light. We describe reactions as either exothermic or endothermic, depending on whether energy is absorbed or given out.

EXCHERMIC REACTIONS

An exothermic reaction is one during which heat is given out, causing a temperature rise in the surroundings.

Examples

- When sodium carbonate is dissolved in a beaker of water, the temperature of the solution rises from 28 °C to 40 °C.
- When methane is burnt, heat energy is evolved and the temperature of the surroundings rises.
- When acids react with alkalis, neutralisation takes place with the evolution of heat. The temperature of the solution formed rises.

ENDOTHERMIC REACTIONS

An endothermic reaction is one where heat is taken in, causing a temperature drop in the surroundings.

Examples

- When ammonium chloride is dissolved in a beaker of water, the temperature of the solution drops from 28 °C to 22 °C.
- Heat energy must be supplied during the thermal decomposition of calcium carbonate.
- Light energy must be absorbed before photosynthesis by plants can take place.

Which of the following statements about endothermic reactions are carrect?

- 1. Energy is released to the surroundings.
- 2. Energy is absorbed from the surroundings.
- 3. The temperature of the surroundings rises.
- 4. The temperature of the surroundings falls.
- A 1 and 2

C 1 and 4

- B 2 and 3
- D 2 and 4

EXAMTIP IN

When trying to decide if a reaction is exothermic or endothermic, always consider the temperature change in the surroundings. If a reaction is exothermic, heat energy will be given out to the surroundings and the surrounding temperature will rise while the total energy content of the reactants will drop. Similarly, if a reaction is endethermic heat energy will be absorbed from the surroundings and cause the surrounding temperature to drop, while the total energy centent of the reactants will rise.

[Ans: D] Since heat is absorbed from the surroundings, the temperature of the surroundings will fall.

ENTINETY CHANGES

Energy level diagrams (Fig 11 and 12) can be drawn to show the changes in energy content of the reactants and products involved in reactions. Enthalpy values or H values are used in energy diagrams.

Hoscians = total energy content of the reactants

H_{products} = total energy content of the products

ΔH = change in energy during reaction = H_{products} - H_{modulus}

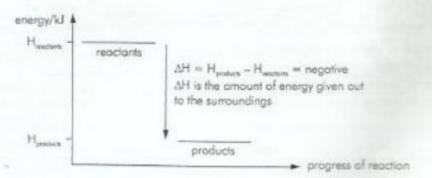


Fig 11: Energy level diagram for exothermic reactions

For exothermic reactions, the enthalpy change is negative, i.e. ΔH = negative. This amount of energy is given out to the surroundings during reaction.

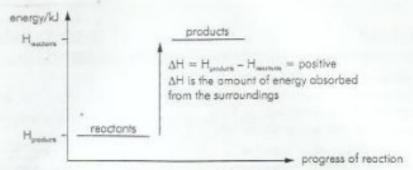


Fig 12: Energy level diagram for endothermic reactions

For endothermic reactions, the enthalpy change is positive, i.e. $\Delta H = positive$. This amount of energy is absorbed from the surroundings during reaction.

BOND BREAUNG AND BOND FORMATION

In a chemical reaction, bonds between reactants must be broken so that the atoms can rearrange themselves to form products. Heat energy must be taken in by the reactants for bond breaking. Hence, bond breaking must be endothermic.

Example: $H - H \rightarrow H + H$ $\Delta H = positive$

Energy is needed to break the covalent band between the two hydrogen atoms in the hydrogen molecule.

After all the bonds in the reactants are broken, the atoms will form new bonds to give the products of the reaction. Heat energy will be released when these new bonds are formed. Bond making is thus exothermic.

Example: $O + O \rightarrow O=O$ $\Delta H = negative$

Energy is released when the two oxygen atoms form a double covalent bond.

STOP AND THINK

Which process is endothermic?

A The formation of a hydrogen-chlorine bond.

B The formation of rust.

C The formation of water from ice.

D The formation of water from oxygen and hydrogen.

EXAMTIP ISP

The reactions in A, B and D are all exothermic reactions.

[Ans: C] Heat energy must be supplied to melt the ice.

STOP AND THINK

Three equations are given.

Equation 1 : H, → 2H

Equation 2 : 2H + O → H₂O

Equation 3 : HCl -> H + Cl

Which equations represent endothermic processes?

A 1 and 2

B 1 and 3

C 2 and 3

D 3 only

EXAMTIP IS

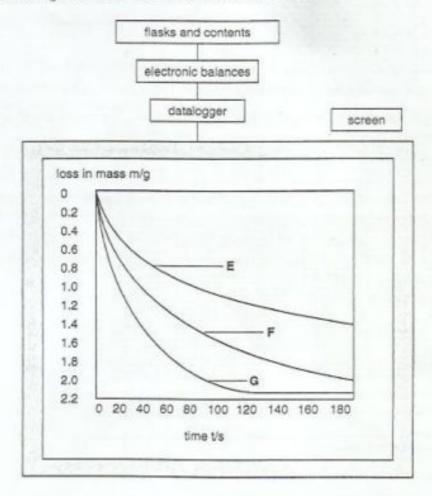
Equation 2 is an exothermic reaction as the hydrogen and exygen atoms will form water, a covalent compound with the structure H - O - H.

[Ans: 8] Heat energy must be supplied break the covalent bends in H - H and H - Cl.

SAMPLE QUESTIONS

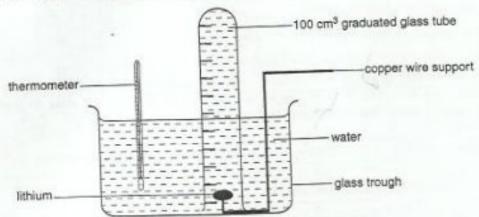
- 1. Three samples of calcium carbonate are placed in flasks for an investigation.
 - In flask E is 5 g of calcium carbonate large lumps.
 - In flask F is 5 g of calcium carbonate medium sized lumps.
 - In flask G is 5 g of calcium carbonate small lumps.

The same volume, an excess, of dilute hydrochloric acid is added to each flask. The flasks are placed on three electronic balances. A data logger is used to plot the loss of mass of the flasks and their contents against time. The results are shown below.



- (a) (i) Why do the three flasks and their contents lose mass?
 - (ii) How do the rates of reaction change with time?
- (b) In which flask is the reaction fastest at time t = 20 s?
- (c) (i) After how long does the reaction in flask G stop?
 - (ii) Why does this reaction stop?
- (d) (i) Sketch on the figure the curve you would expect if 5 g of powdered calcium carbonate is used instead of 5 g of lumps of calcium carbonate. Label this curve H.
 - (ii) Use the particle theory to explain why the size of the lumps of calcium carbonate affects the rate of reaction.
- (e) What name is given to a reaction in which heat is given out?

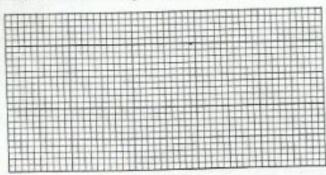
- The metal lithium reacts with air and water.
 - (a) Suggest how lithium should be stored.
 - (b) A student reacted lithium with water using the apparatus as shown.



The student measured the valume of gas at intervals of 30 seconds. The results are shown below.

time / s	0	30	60	90	120	150
volume / cm ³	0	40	60	74	86	96

Plot a graph of these results on the grid below. Use the vertical axis to plot volume.



(c) Tick a box to show when the rate of reaction was greatest.

at time / s	5	35	65	95	125
greatest rate of reaction					

- (d) How can the student find the time taken for the reaction to stop?
- (e) The temperature of the water increased during the reaction. How would you classify the reaction?
- (f) (i) Name the gas liberated in this reaction.
 - (ii) Describe a test to confirm the presence of this gas.
- (g) (i) Universal indicator is added to the solution in the trough at the end of the experiment. What colour will you see?
 - (ii) What ions, present in this solution, cause this change?

3. Three of the reactions below involve both oxidation and reduction. Tick these in the appropriate box.

CaCO ₂ (s)	->	CaO (s) + CO ₂ (g)	
N ₂ (g) + 3H ₂ (g)	\rightarrow	2NH, (g)	
$Zn(s) + Cu^{2+}(aq)$.	-	Zn ²⁺ (aq) + Cu (s)	
NH ₄ CI (s)	\rightarrow	range for the standard for	
S(s) + O ₂ (g)	\rightarrow	SO ₂ (g)	
H ₂ O (I)	\rightarrow	H ₂ O (g)	
2NaOH (aa) + CuSO, (aa)	->	Cu(OH), (s) + Na.5O (aa)	

4. (a) Oxidation can be described as removal of electrons.

Look at the reactions below and state whether the substance underlined has been oxidised, reduced or has been neither oxidised nor reduced.

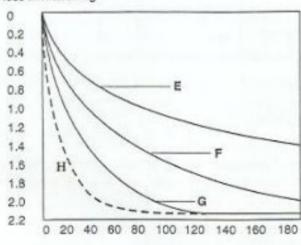
(i)
$$2K(s) + Br_2(g) \rightarrow 2K^+Br^-(s)$$

(ii) $2A^{3+}(s) + 3Mg^{2+}(s) \rightarrow 3Mg^{2+}(s) + 2AI(s)$
(iii) $H^+(aq) + OH^-(aq) \rightarrow H_2O(I)$

- (b) When magnesium metal is added to a solution of lead(II) nitrate a redox reaction takes place. Grey crystals of lead and a colourless solution containing magnesium ions are formed.
 - (i) Write the simplest ionic equation for the redox reaction.
 - (ii) What is the oxidising agent in the reaction?

SOLUTIONS AND EXPLANATIONS

- (a) (i) The reaction between calcium carbonate and dilute hydrochloric acid produces carbon diaxide gas, which escapes into the surroundings.
 - (ii) The reaction rate is the fastest at the beginning. As the reaction proceeds, the reaction rate decreases until the reaction stops.
 - (b) Flask G
 - (c) (i) The reaction stops after 130 seconds
 - (ii) All the calcium carbonate has been used up.
 - (d) (i) loss in mass m/g

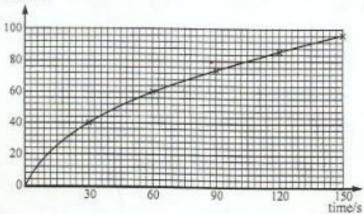


- time t/s
- (ii) With smaller lumps of calcium carbonate, the total surface area increases, allowing for more effective collisions between the calcium carbonate and the hydrochloric acid molecules to occur. This increases the speed of reaction.
- (e) Exothermic reaction.

EXAMTIP IS

When a graph is given, the student is expected to extract data from it to answer the questions that follow. Hence it is not sufficient for the student to give an answer to (a)(ii) as 'speed of reaction decreases'. The student is expected to divide the graph into sections and explain the changes in the curve occordingly.

- 2. (a) Lithium can be stored under ail.
 - (b) volume/cm3



(c)	at time / s	5	35	65	95	125
	greatest rate of reaction	1				

- (d) The time taken for the reaction to stop is the time period from the start of the experiment to the time when there is no more effervescence and the volume of hydrogen gas collected remains constant.
- (e) Exothermic reaction.
- (f) (i) Hydrogen gas
 - (ii) It will cause a lighted splint to extinguish with a pop sound.
- (g) (i) A blue solution will be formed.
 - (ii) Hydroxide ions.

EXAMTIP IS

Lithium, a very reactive metal, belongs to Group I in the Periodic Table. It reacts with water to produce an alkaline lithium hydroxide solution and hydrogen gas.

2Li + 2H,O → 2LiOH + H,

Universal indicator will turn blue or violet in the presence of a strong alkaline solution, and red in the presence of a strong acid such as dilute hydrochloric acid.

3.	CoCO ₃ (s)	\rightarrow	$CaO(s) + CO_2(g)$	
	$N_2(g) + 3H_2(g)$	\rightarrow	2NH ₃ (g)	
	$Zn (s) + Cu^{2+} (aq)$	\rightarrow	Zn2+ (aq) + Cu (s)	
	NH _e Cl (s)	\rightarrow	NH ₃ (g) + HCl (g)	
	S (s) + O ₂ (g)	\rightarrow	SO ₂ (g)	1
	H ₂ O (I)	\rightarrow	H ₂ O (g)	
	2NaOH (ac) + CuSO (aa)	-	Cu(OH) (s) + No SO (oa)	

EXAMTIP 53

 $N_{+}(g) + 3H_{+}(g) \rightarrow 2NH_{+}(g)$

N, is reduced to NH, since the oxidation state of nitrogen has decreased from 0 to -3.

 H_2 is exidised to NH, since the exidation state of hydrogen has increased from 0 to *1. $Zn(s) + Cu^{2s}(aq) \rightarrow Zn^{2s}(aq) + Cu(s)$

Zn is oxidised to Zn^{α} since the oxidation state of zinc has increased from 0 to *2. Cu' is reduced to Cu since the oxidation state of copper has decreased from *2 to 0. $S(s) + O_{\alpha}(g) \rightarrow SO_{\alpha}(g)$

5 is oxidised to SO, since it has goined oxygen.

Thus O, must be reduced since exidation and reduction must occur together.

4. (a) (i) Oxidised

- (ii) Reduced
- (iii) Neither
- (b) (i) Mg + Pb²⁺ \rightarrow Mg²⁺ + Pb
 - (ii) Pb2+ is the oxidising agent.

EXAMTIP IS

In (a)(i), K has last one electron to form K': it is oxidised. In (a)(ii), All has gained 3 electrons to form Al. it is reduced. There is no change in oxidation state for H' in (a)(iii), hence no redox reaction has taken place.

In (b)(ii), Pb^b has gained electrons to form Pb. It is reduced in the reaction and is thus the oxidising agent.



ACIDS, BASES AND SALTS

LEARNING OBJECTIVES

Candidates should be able to:

6.1 Characteristic Properties of Acids and Bases

- describe the meanings of the terms acid and alkali in terms of the ions they contain or produce in aqueous solution and their effects on Universal Indicator paper
- describe how to test hydrogen ion concentration and hence relative acidity using Universal indicator paper and the pH scale
- describe the characteristic properties of acids as in reacts with metals, bases and carbonates
- describe the importance of controlling the pH in soils and how excess acidity can be treated using calcium hydroxide
- describe the characteristic properties of bases as in reactions with acids and with ammonium salts
- I classify oxides as acidic, basic or amphoteric, based on metallic/non-metallic character

6.2 Preparation of Salts

- describe the techniques used in the preparation, separation and purification of salts as examples of some of the techniques specified in section 1.2 – methods for preparation should include precipitation and titration, together with reactions of acids with metals, insoluble bases and insoluble carbonates
- suggest a method of preparing a given salt from suitable starting materials, given appropriate information

USEFUL WEBSITES

ш	http://www.chemistry.co.nz/acids_and_bases.htm	
	http://www.quia.com/dir/chem/	

- ☐ http://www.chem4kids.com/files/react_acidbase.html
- □ http://www.krysstal.com/acidbase.html
- □ http://www.bbc.co.uk/scotland/education/bitesize/standard/chemistry/acids/index.shtml
- http://www.sciencebyjones.com/acids_bases_salts.htm
- http://www.school-for-champions.com/science/chemsalts.htm
- ☐ http://www.angelo.edu/faculty/kboudrea/demos.htm

OVERVIEW OF THE TOPIC

ACIDS, BASES AND SALTS ** key areas

6.1	CHARACTERISTIC PROPERTIES OF ACIDS AND BASES	& Acids	Production of H ⁺ ions
			Physical properties of acids
			Reaction of acids with metals
			Reaction of acids with bases
			Reaction of acids with carbonates
		BASES	Bases and alkalis
			Physical properties of alkalis
			Reaction of bases with ammonium salts
		⊚ ≯H	• pH scale
			Effect of acids and alkalis on universal indicator paper
			Control of pH in agriculture
		€ Oxides	Classification of oxides
			Properties of different types of oxides
6.2	PREPARATION OF SALTS	Solubility Rules	
		PREPARATION OF INSOLUBLE SALTS	Precipitation
		PREPARATION OF SOLUBLE SALTS	Action of acid on an insoluble metal/base/carbonate
			Titration

CHARACTERISTIC PROPERTIES OF ACIDS AND BASES



(a) Production of H+ Ions

Acids are substances that will dissolve in water and undergo ionisation to form hydrogen ions. The table below shows some common acids found in the laboratory and the ions they contain.

Name of Acid	lons present
Hydrochloric ocid, HCI	H+, Cl-
Sulphuric acid, H ₂ SO ₄	H+, SO ₄ ² -
Nitric acid, HNO ₃	H+, NO ₃ -
Ethanoic acid, CH ₃ COOH	H+, CH,COO-

Table 1: Common Acids

The properties and reactions of an acid are due to the presence of hydrogen ions, H*. Acids have acidic properties only when they are dissolved in water.

EXAMTED IS

Note that HCl in gaseous form is called hydrogen chloride. If it is dissolved in water, it will undergo ionisation to form a solution called hydrochloric acid.

(b) Physical Properties of Acids

- Acids have a sour taste. For example, vinegar and lemon are both sour because they contain ethanoic acid and citric acid respectively.
- Acids change the calour of indicators. For example, blue litmus paper turns red when dipped in acid.
- Acids have pH values less than 7.

(c) Reaction of Acids with Metals

Acids react with metals above hydrogen in the reactivity series to give salts and hydrogen gas.

Example 1: Hydrochloric acid + magnesium → magnesium chloride + hydrogen gas 2HCl (aq) + Mg (s) → MgCl, (aq) + H_s (g)

We can test for hydrogen gas using a lighted splint. The flame extinguishes with a 'pop' sound.

EXAMTIP TOP

The metal reactivity series lists metals according to their reactivity. The more reactive metals are placed at the top of the list, while the less reactive metals are placed near the bottom of the list, Hydrogen (a non-metal) is inserted into the list as a reference point - metals above hydrogen will react with dilute acids to produce hydrogen, while metals below hydrogen will not. (See topic B)



(d) Reaction of Acids with Bases

Acids react with bases to give salts and water only. Bases are usually metal axides or metal hydroxides.

Example 2: Acid with Metal Oxide

Hydrochloric acid + zinc oxide → zinc chloride + water

2HCl (aq) + ZnO (s) → ZnCl₂ + H₂O

Example 3: Acid with Metal Hydroxide

Hydrochloric acid + sodium hydroxide → sodium chloride + water

HCl (aq) + NaOH (aq) → NaCl (aq) + H₂O (aq)

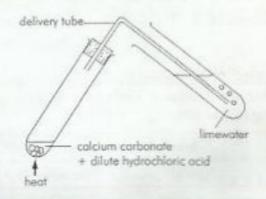
[11] This reaction is also called neutralisation.

(e) Reaction of Acids with Carbonates

Acids react with carbonates to give salts, carbon dioxide gas and water.

Example 4: Calcium carbonate + hydrochloric acid \rightarrow calcium chloride + carbon dioxide + water $CaCO_3$ (s) + 2HCl (aq) \rightarrow CaCl₂ (aq) + CO_2 (g) + H_2O (l)

We can test for carbon dioxide gas using limewater (calcium hydroxide solution). Carbon dioxide causes the limewater to turn chalky.



STOP AND THINK

Dilute hydrochloric acid, dilute nitric acid and dilute sulphuric acid all react with

- A aqueous barium chloride to give a precipitate.
- B copper to give hydrogen.
- C aqueous silver nitrate to give a precipitate.
- D aqueous sodium carbonate to give carbon diaxide.
- E aqueous sodium hydroxide to give hydrogen.

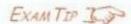
EXAMTED TOP

In A, only sulphuric acid will react with barium chloride to give a white precipitate of barium chloride. In B, copper is below hydrogen in the reactivity series, and so will not react with any of the three acids to give hydrogen gas. In C, hydrochloric acid is the only acid that will react with silver nitrate to give a white precipitate of silver chloride. D is wrong because acids will react with bases to give salt and water only.

[Ans:D] All three acids will react with sodium carbonate to give carbon dioxide gas and water. The salts formed are sodium chloride (with hydrochlaric acid), sodium nitrate (with nitric acid) and sodium sulphate (with sulphuric acid).

Dilute sulphuric acid reacts with copper(II) oxide and with copper(II) carbonate. In which ways are these two reactions alike?

- A A gas is produced.
- B An insoluble salt- is precipitated.
- C Sulphuric acid reacts as an oxidising agent.
- D Water is a product.



Acid-base and acid-carbonate reactions are not redox reactions. In the question above, there is no change in oxidation state for any of the elements involved in both reactions. [Ans: D] Dilute sulphuric acid reacts with copper(II) oxide to form aqueous copper(II) sulphate and water. When it reacts with copper(II) carbonate, aqueous copper(II) sulphate, carbon dioxide and water are formed.



(a) Bases and Alkalis

A base is a substance that reacts with an acid to give a salt and water only.
Example 5: Hydrochloric acid + sodium hydroxide → sodium chloride + water

 $HCI(aq) + NaOH(aq) \rightarrow NaCI(aq) + H_2O(aq)$

Bases are usually metal oxides or metal hydroxides. Many bases are insoluble in water. Bases that can dissolve in water form solutions called alkalis. Table 2 below lists some common bases and alkalis.

Insoluble bases		Soluble bases	
Name	Formula	Name	Formula
magnesium oxide	MgO	sodium hydroxide	NaOH
copper(II) oxide	CuO	calaium hydroxide	Ca(OH),
lead(II) oxide	РЬО	ammonium hydroxide (or aqueous ammonia)	NH,OH

Table 2: Common bases and alkalis

Hydroxide ions, OH, are produced when bases dissolve in water to form alkalis.

Example 6: Sodium hydroxide, NaOH

NaOH (aq) -> Na* (aq) + OH (aq)

The ability of alkalis to neutralise acids is due to the presence of these hydroxide ions.

What is the ionic equation for the neutralisation of aqueous sodium hydroxide with dilute nitric acid?

A
$$H^+ + OH^- \rightarrow H_0O$$

Students must be able to recognise the ionic equation: $H^* + OH \rightarrow H_*O$ for acid-alkali neutralisation.

[Ans: A] The ionic equation shows only the ions involved in the reaction. In neutrolisation, water is formed from the H' ions that are responsible for acidic reactions and the OH ions that are responsible for alkali reactions.

(b) Physical Properties of Alkalis

- Alkalis feel slippery.
- Edible alkalis have a bitter taste.
- Alkalis change the colour of indicators. For example, red litmus paper will turn blue when dipped into an alkaline solution.
- The pH of alkaline solutions is greater than 7.

(c) Reaction of Bases with Ammonium Salts

Alkalis react with ammonium salts to produce salts, ammonia gas and water.

Example 7: Sodium hydroxide + ammonium chloride

→ sodium chloride + ammonia + water

NaOH (aq) + NH₄Cl (aq) \rightarrow NaCl (aq) + NH₃ (g) + H₂O (l)

The ammonia gas evolved is pungent, colourless and turns damp red litmus paper blue.



STOP AND THINK

A sample of ammonium sulphate is warmed with aqueous sodium hydroxide. The gas produced

- A ignites with a 'pop'.
- B relights a glowing splint,
- C turns damp red litmus paper blue.
- D turns limewater cloudy.



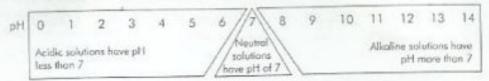
Students must differentiate between the ammonium ion test and the nitrate test. In both tests, ammonia gas is produced. However, in the ammonium ion test, only aqueous sodium hydroxide is added before warming: the nitrate test requires one to add sodium hydroxide AND aluminium foil before warming to give ammonia gas.

[Ans: C] Ammonium salts react with alkalis on warming to produce ammonia gas. Ammonia gas is an alkaline gas that will turn damp red litmus paper blue.



(a) pH Scale

pH measures how acidic or how alkaline a solution in water is. It is a measure of the concentration of H+ ions present in the solution.



- The pH scale ranges from 0 to 14.
- A pH value of 7 indicates that the solution is neutral, i.e. the number of H+ and OH+ ions in the solution are the same.
- A pH value of less than 7 indicates that the solution is acidic. The solution has a greater concentration of H⁺ ions than OH⁻ ions.
- A pH value of more than 7 indicates that the solution is alkaline. The solution has a lower concentration of H+ ions than OH- ions.



An aqueous solution will always have both H' and OH ions present irregardless of its pH value. This is because water itself undergoes ionisation to give H' and OH.

(b) Effect of Acids and Alkalis on Universal Indicator paper

We can use Universal Indicator to measure the pH of a solution. Fig 1 below shows the colour change in the Universal Indicator with pH.

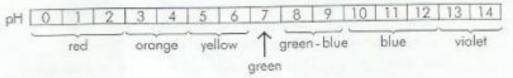


Fig 1: Colour change in Universal Indicator with pH

The Universal Indicator is not very accurate in determining the pH of a solution. It is also not possible to measure the pH of coloured solutions because the colour change in the indicator cannot be seen. For accurate measurements, a pH meter is used.



When we measure pH, we are measuring the concentration of H+ ions in aqueous solution. It is not possible to measure the pH of an insoluble substance.

STOP AND THINK

The colour of Universal Indicator in each of the four solutions is shown in the table. Which solution is likely to be an acid of pH 2-3?

Solution	Colour of Universal Indicator
Α	blue
В	green
C	red
D	violet

EXAMTIP IN

The range of colours in different solutions of pH for the universal indicator approximates the rainbow colours - red, arange, yellow, green, blue, indigo and vialet. Taking pH 7 (neutral) to be green, colours to the left of green in the rainbow indicate acidic solutions, while colours to the right indicate alkaline solutions.

[Ans. C] Solution 8 has pH 7, while solutions A and D are alkaline, i.e. their pH is greater

than 7.

(c) Control of pH in agriculture

- Most plants need a soil pH of 6.5 to 7.5 to grow well. If the ground is too acidic, slaked lime (solid calcium hydroxide) can be added to neutralise the acid. This process is called liming the soil.
- Slaked lime is chosen because
 - 1 It is cheap and easily available.
 - Slaked lime is sparingly soluble in water. Once the acid is neutralised, the excess base will remain as solid in the soil. It will not dissolve in water to make the soil too alkaline.

EXAMTIP 30

The soil is often too acidic because of acid rain.

Other possible chemicals that can be used to neutralise excess acid are calcium carbonate (limestone) and calcium exide (lime).

STOP AND THINK

In an accident at a factory, some nitric acid was spilt. Which substance, when added in excess, would neutralise the acid without leaving an alkaline solution?

A aqueous ammonia

C calcium carbonate

B aqueous sodium hydroxide

D water

EXAMTIP JO

Both aqueous ammonia and aqueous sodium hydroxide are alkalis that will neutralise nitric acid. However, the person spraying the solution (e.g. sodium hydroxide solution) will not know when enough alkali has been added to neutralise the acid since the products of neutralisation appears as a colourless solution (sodium nitrate and water). Excess alkali will cause the ground to become alkaline.

[Ans C] Calcium carbonate is insoluble in water. When all the nitric acid is neutralised, the axcess calcium carbonate will remain in solid form on the ground to be swept away.



(a) Classification of oxides

- Oxides are formed when substances burn in oxygen gas. Oxides have acidic, basic or amphateric character, depending on which type of oxide they belong to.
- Table 3 lists 3 types of oxides.

Type of Oxides	Examples
Acidic oxide	nítrogen dioxide, NO ₂ sulphur dioxide, SO ₂ carbon dioxide, CO ₂
Basic oxide	sodium oxide, Na ₂ O magnesium oxide, MgO calaium oxide, CaO
Amphoteric oxide	zinc oxide, ZnO aluminium oxide, Al ₂ O ₃ lead(II) oxide, PbO

Table 3: Types of oxides

(b) Properties of different types of oxides

Acidic oxides are usually oxides of non-metals. They form acids when dissolved in water.

Example 8: Carbon dioxide + water \rightarrow carbonic acid CO₂ + H₂O \rightarrow H₂CO₃

EXAMTIP JOS

Natural rain has a pH slightly lower than 7. Carbon dioxide in the air will dissolve in rainwater to produce a weakly acidic solution of carbonic acid.

Basic oxides are oxides of metals. They react with acids to produce salt and water only.

Example 9: Copper(II) oxide + sulphuric acid → copper(II) sulphate + water

CuO + H₂SO, → CuSO, + H₂O

Neutralisation takes place. Basic oxides that dissolve in water form solutions called alkalis.

Some oxides of metals behave as acidic or basic oxides. When they react with acids, they behave as basic oxides; when they react with alkalis, they behave as acidic oxides. These oxides are known as amphoteric oxides.

Example 10: Aluminium oxide + hydrochloric acid \rightarrow aluminium chloride + water Al_2O_3 + 6HCI \rightarrow $2AICI_3$ + H_2O

In this case, aluminium oxide is acting as a base.

Aluminium oxide + sodium hydroxide → sodium aluminate + water In this case, aluminium oxide is acting as an acid.

EXAMTIP 100

There is a fourth class of oxides called neutral oxides. They do not dissolve in water to form acids nor do they react with bases to form salts. Examples of such oxides are carbon monoxide and water.

STOP AND THINK

Which element forms an acidic oxide?

A aluminium C phosphorus

magnesium D sodium

EXAMTER 300

As a general rule, metals form basic oxides, except for aluminium, lead and zinc, which form amphotoric oxides.

[Ans: C] Generally, non-metals combine with oxygen to form acidic oxides. The oxide of phosphorus is acidic and will dissolve in water to form an acidic solution.

PREPARATION OF SALTS

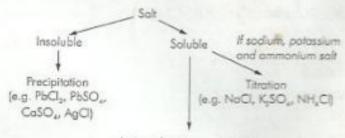


- The method chosen to prepare a salt depends on its solubility.
- Table 4 below lists the solubilities of some classes of compounds. Note that all sodium, potassium, ammonium and nitrate compounds are soluble.

Solubility of Compounds	Solubility	
Carbonates	All insoluble except for carbonates of sodium, potassium and ammonium	
Oxides	All insoluble except for oxides of sodium, potassium and calcium	
Sulphates	All soluble except for sulphates of calaium (sparingly soluble) lead, barium and silver	
Chlorides	All soluble except for chlorides of silver and lead	
Hydroxides	All insoluble except for hydroxides of sodium, potassium, ammonium, calcium (sparingly soluble) and magnesium (very sparingly soluble)	

Table 4: Solubility of salts

The selection of salt preparation method is summarised in Fig 2 below. Precipitation is carried out if an insoluble salt is required. If a soluble salt is needed, it is prepared by either the action of an acid on a metal/metal oxide/carbonate or by titration.



Action of acid on metal oxide
Action of acid on metal carbonate
Action of acid on metal (not for Cu salts)
(e.g. CuSO₄, MgCl₂, Ca(NO₃)₂)

Fig 2: Selecting the correct preparation method

PREPARATION OF INSCRUBE SALTS USING PRECIPITATION

- The preparation of lead(II) iodide is used to illustrate the procedure.
 - Choose 2 starting solutions. One must contain the positive ion of the insoluble salt, the other must contain the negative ion. Mix the solutions to precipitate out the salt.

 $Pb(NO_3)_2 (aq) + 2Nal (aq) \rightarrow Pbl_2 (s) + 2NaNO_3 (aq)$

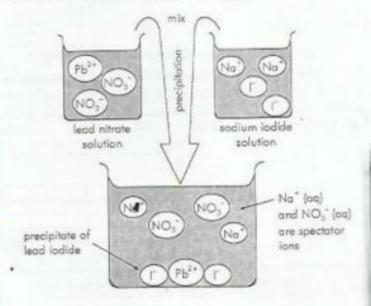
STEP © Filter and wash the precipitate with distilled water. Dry between pieces of filter paper.

The ionic equation for this reaction is $Pb^{2+}(aq) + 2l^{-}(aq) \rightarrow Pbl_{2}(s)$

Other salts prepared using precipitation include silver chloride, barium sulphate and lead chloride.

EXAMTIP TOP

The reactants involved in a precipitation reaction must be in solution form because the ions must be able to move and interact with one another when the reactants are mixed together. When the ions in the insoluble salt encounter each other, they will attract each other to form a solid that will sink to the bottom of the container and be collected as the precipitate.



To ensure that the reactants chosen are soluble in water, the easiest choice is to select sodium and nitrate compounds which are all soluble. For example, if silver chloride is the salt required, then we can choose to mix sodium chloride solution and silver nitrate solution together to form the salt. If barium sulphate is needed, then barium nitrate solution and sodium sulphate solution can be used.

STOP AND THINK

Lead(II) sulphate is insoluble in water. What should be added to dilute sulphuric acid to prepare lead(II) sulphate?

- A lead metal
- B powdered lead(II) carbonate
- C powdered lead(II) chloride
- D aqueous lead(II) nitrate
- E powdered lead(II) oxide

EXAMTIP TOP

Lead, lead(II) carbonate and lead(II) oxide cannot be used to prepare lead(II) sulphate because they are all insoluble in water. When sulphuric acid is added to any one of these solid reactants, the reaction will stop once they are coated with a layer of insoluble lead(II) sulphate, leading to very poor yields of the salt.

[Ans: D] Solid lead(II) sulphate and nitric acid will be produced in this reaction.

PREPARATION OF SOLUBLE SALTS

(a) Action of Acid on an Insoluble Metal/Base/Carbonate

- This method is used for preparation of soluble salts when a suitable insoluble starting material can be found. This starting material can be a metal, a metal oxide or a carbonate.
- The preparation of copper(II) sulphate is used to illustrate the procedure. Fig 3 illustrates the whole process.

Starting materials: copper(II) oxide and dilute sulphuric acid.

$$CuO(s) + H_sO_s(aq) \rightarrow CuSO_s(aq) + H_sO(l)$$

- Warm dilute sulphuric ocid in a beaker but do not boil. The acid is warmed to increase the speed of reaction between the reactants.
- Add copper(II) axide pawder to the acid with stirring until no more can react. This ensures that all the acid has been used up.
- STEP 3 The mixture contains copper(II) sulphate solution and excess copper(II) oxide. Filter to remove copper(II) oxide.
- STEP
 The filtrate is heated in an evaporating dish until a thin crust of crystals forms on the surface of the liquid. It is then left to coal for more crystals to form.
- STEP © The crystals are then filtered and dried between filter paper before collection.

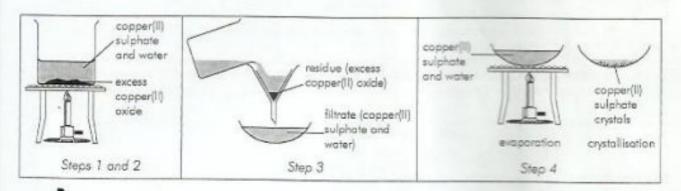


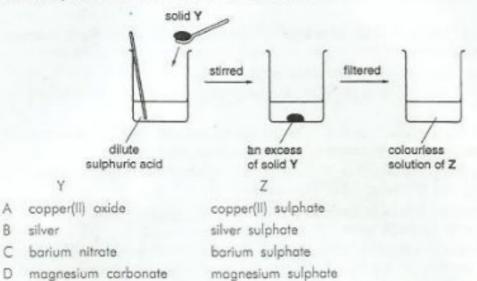
Fig 3: Preparing a soluble salt by action of acid on a metal/base/metal carbonate.

EXAMTIP 100

If a metal carbonate is used to prepare a salt using this method, there will be effervescence of carbon diaxide gas as the metal carbonate is added to the acid in Step 2. When there is no more effervescence, all the acid has been used up and we may proceed to Step 3.

STOP AND THINK

The diagrams below show how a colourless solution of Z is produced from the reaction between sulphuric acid and a solid Y. What are Y and Z?



EXAMTER TOP

In A. copper(II) sulphate can be made using the method described in the question, but a blue filtrate will be obtained.

[Ans: D] Magnesium carbonate is insoluble in water. Excess solid magnesium carbonate will need to be filtered away after all the sulphuric acid has reacted. The filtrate is a colourless solution of magnesium sulphate.

(b) Titration

- This method is used when it is not possible to find a suitable insoluble starting material like a metal, a metal oxide or a carbonate that can be easily filtered off at the end of the maction. All the starting materials and the salt product are solutions.
- The preparation of sodium nitrate is used to illustrate the procedure. Fig 4 and 5 illustrate the process.

Starting materials: aqueous sodium hydroxide and dilute nitric acid

$$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

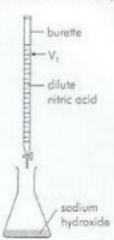
In a titration, an indicator is needed to show the endpoint of the reaction. The endpoint is the volume of one reactant needed to exactly neutralise a given volume of the other reactant. A common indicator used in the laboratory is the screened methyl orange. The calour changes for this indicator are shown below.

acidic solution | alkaline solution | Indicator is green

Endpoint Indicator is grey or colourless A preliminary titration is carried out to find the endpoint with the help of an indicator.



Using a pipette, measure 25.0 cm³ of sadium hydroxide into a conical flask. Add two drops of an indicator to the alkali.



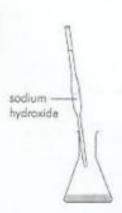
Add dilute nitric acid from a burette to sodium hydroxide until the solution just turns colour. At this point, just enough acid is added to neutralise the alkali. All the alkali has reacted. The resulting mixture contains only sodium nitrate and water.



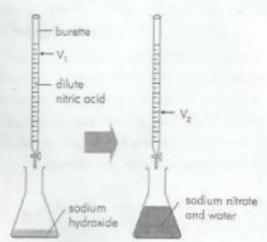
Prom the titration result, we can know the exact volume of nitric acid needed to react with 25.0 cm³ of sodium hydroxide. Volume of nitric acid, V_x = V₂ - V₁

Fig 4: Preliminary titration

A second titration is carried out without the indicator. The exact volume obtained from the preliminary titration is used.



 Using a pipette, measure 25.0 cm³ of sodium hydroxide into a conical flask. This time, no indicator is added.



- Add V_a cm³ of dilute nitric acid from a burette to the sodium hydroxide. Sodium nitrate is formed in the solution.
- © Evaporate the mixture to obtain a saturated solution. Cool the saturated solution to obtain sodium nitrate crystals. The mixture is filtered to obtain pure sodium nitrate crystals.

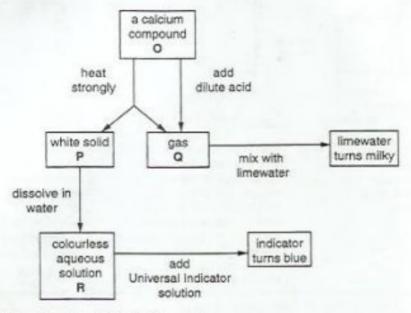
Fig 5 The actual titration to prepare the salt

EXAMTIP IS

The salt solution obtained in the preliminary titration is thrown away because it is contaminated by the indicator. This is why a second titration without the indicator is carried out.

SAMPLE QUESTIONS

- (a) Lime is an alkaline substance. Give two uses of lime that depend on this property. One of the
 uses must be in farming.
 - (b) The figure shows some of the properties of a calcium compound, O. The letters are not the chemical symbols of the substances.



- (i) Suggest the identity of O, P, Q and R.
- (ii) Name an acid that could be used to liberate gas Q from solid O.
- (iii) Write an equation, including state symbols, for any one of the reactions shown.
- 2. (a) How can ammonia gas be made from ammonium sulphate?
 - (b) When ammonia is dissolved in water, aqueous ammonia results.
 - (i) How could you show that aqueous ammonia is alkaline?
 - (ii) What causes aqueous ammonia to be alkaline?
 - (c) An excess of dilute hydrochloric acid is added to aqueous ammonia drap by drap.
 - (i) Describe what happens to the pH of the mixture.
 - (ii) State the formula of the salt formed.
- 3. (a) This table is about the acidity or alkalinity of three liquids.
 - (i) Complete the table.

Liquid	pH	colour with universal indicator
hydrochloric acid		red
water	7	
aqueous sodium hydroxide		blue

- (ii) Briefly explain, in terms of the ions in solution, the reason for the differences in acidity and alkalinity of these three liquids.
- (b) A farmer decided to neutralise and fertilise his sail by adding at the same time both calcium oxide and an ammonium salt. Why was he wasting his time and money?

No.6 Acids, Bases And Salts



4. Here is a list of oxides.

aluminium oxide calcium oxide carbon monoxide sodium oxide

carbon dioxide sulphur dioxide

Use any oxide once, more than once, or not at all. Choose from the list an oxide which is

- (a) commonly called quicklime
- (b) a basic oxide
- (c) an amphoteric oxide
- (d) a cause of acid rain
- (e) a poison found in car exhaust fumes
- (f) produced when one of the others burn
- Most acids have similar chemical reactions. Describe three reactions typical of acids. Write a chemical equation for one of these reactions.
- Copper(II) sulphate crystals may be prepared in the way described below. Read the description and answer the questions that follow.

A small amount of solid copper(II) carbonate was added to dilute sulphuric acid in a beaker. The mixture was stirred until all the copper(II) carbonate had reacted. This was repeated until no more copper(II) carbonate would react. The mixture was filtered into an evaporating basin which was warmed until a saturated solution was abtained and then allowed to cool. The crystals produced were filtered off, washed and dried before being weighed.

- (a) Describe all that you would see from the time the first sample of copper(II) carbonate was added to the acid until no more would react. Explain how you would know that all the copper(II) carbonate had reacted.
- (b) What is meant by the term saturated solution? Describe one test you would make on the solution of copper(II) sulphate to see that it was saturated before it was allowed to coal.
- (c) Describe how you would wash and dry the crystals.

SOLUTIONS AND EXPLANATIONS

- (a) Lime is mixed into soil to neutralise excess acidity. Wet lime is also used to remove sulphur dioxide from exhaust gases of coal burning power stations.
 - (b) (i) O is calcium carbonate

P is calcium oxide

Q is carbon dioxide

R is calcium hydroxide

(ii) Hydrochloric acid

(iii) $CaCO_1(s) \rightarrow CaO(s) + CO_1(g)$

EXAMTIP TOP

Most carbonates decompose on heating to form metal exides and carbon diexide gas. Since calcium exide is a basic exide, it will dissolve in water to form calcium hydroxide, which is alkaline (as shown by the blue colour of the Universal indicator added). Other possible equations that can be written are

 $CaO(s) + H_{*}O(l) \rightarrow Ca(OH)_{*}(aq)$

 $CaCO_{\gamma}(s) + 2HCI_{\gamma}(aq) \rightarrow CaCI_{\gamma}(aq) + CO_{\gamma}(q) + H_{\gamma}O_{\gamma}(1)$

- (a) Add sodium hydroxide or calcium hydroxide to ammonium sulphate and warm the mixture gently.
 A colourless, pungent gas that turns damp red litmus paper blue will be produced. This gas is ammonia gas.
 - (b) (i) The solution will turn universal indicator blue.
 - (ii) The presence of hydroxide ions in aqueous ammonia makes it alkaline.
 - (c) (i) The pH of the mixture will drop from a value around 12 as the hydrochloric acid neutralises aqueous ammonia. The pH will have a pH value of about 2 when excess hydrochloric acid is present.
 - (ii) NH₂Cl

EXAMTIP IS

Another way to show that aqueous ammonia is alkaline is to use red litmus paper. It will turn blue.

3. (a) (i)

Liquid	pH	colour with universal indicator
hydrochloric acid	1	red
water	7	green
aqueous sodium hydroxide	14	blue

- (ii) Hydrochloric acid is acidic because it contains more hydrogen ions than hydroxide ions. Aqueous sodium hydroxide is alkaline because it contains more hydroxide ions than hydrogen ions. Water is neutral as it has equal numbers of hydrogen and hydroxide ions.
- (b) Calcium axide will react with the ammonium salt on warming under the sun to produce ammonia gas which escapes into the atmosphere. In this way, plants will not be able to obtain the nitrogen they need for protein synthesis.

EXAMTIP IS

Besides nitrogen, plants also need phosphorus and potassium for proper growth. These elements are supplied to plants in the form of NPK fertilisers (N for nitrogen, P for phosphorus and K for potassium).

- 4. (a) calcium oxide
 - (b) calcium oxide or sodium oxide
 - (c) aluminium oxide
 - (d) sulphur dioxide
 - (e) carbon monoxide
 - (f) carbon dioxide
- (i) Acids will react with metals above hydrogen in the reactivity series to form salts and hydrogen gas.

- (ii) Acids will react with carbonates to form salts, carbon dioxide and water.
- (iii) Acids will react with bases to form salts and water only.

EXAMTIP TO

The other two equations are: (ii) $CaCO_1 + 2HCI \rightarrow CaCI_2 + CO_2 + H_2O$ and (iii) NaOH + HNO₁ \rightarrow NaNO₂ + H₂O

- (a) Effervescence of a colourless, odourless gas will be observed as the green copper(II) carbonate
 is added to the acid. A blue solution will be formed gradually. When there is no more effervescence,
 the reaction has stopped and no more copper(II) carbonate will react.
 - (b) A saturated solution holds the maximum amount of solute at a given temperature. A few crystals of copper(II) sulphate can be added to the solution. If they dissolve, then the solution is not saturated.
 - (c) The crystals are washed with a little cold distilled water and then dried between filter paper.



The colour of a compound can give a clue as to what ions are present. If copper(II) ions are present in a solution, the solution is usually blue. Iron(II) ions are pale green in solution, while iron(III) ions are pale yellow in solution. This property arises because many transition metals form coloured ions.



THE PERIODIC TABLE

LEARNING OBJECTIVES

Candidates should be able to:

7.1 Periodic Trends

- describe the Periodic Table as an arrangement of the elements in order of increasing proton number (atomic number)
- describe the change from metallic to non-metallic character from left to right across a period of the Periodic Table
- describe the relationship between group number, number of valence electrons and metallic/ non-metallic character
- ☑ predict the properties of elements in Group I, VII and the Transition elements using the Periodic Table
- describe lithium, sodium and potassium in Group I (the alkali metals) as a collection of relatively soft, low density metals showing a trend in melting point and in their reaction with water
- describe chlorine, bromine and iodine in Group VII (the halogens) as a collection of diatomic non-metals showing a trend in colour, state and their displacement reactions with solution of other halide ions

7.2 Group Properties

- describe the elements in Group 0 (the noble gases) as a collection of monoatomic elements that are chemically unreactive and hence important in providing an inert atmosphere (e.g. argon and nean in light bulbs; helium in balloons; argon in the manufacture of steel)
- In describe the lack of reactivity of the noble gases in terms of their electronic structures

USEFUL WEBSITES

	http://www.chem4kids.com/files/elem_intro.html
П	
	http://www.bbc.co.uk/scotland/education/bitesize/standard/chemistry/propertiesofsubstances/index.shtml
	http://web.buddyproject.org/web017/web017/metals.html
	http://www.chemsoc.org/viselements/pages/periodic_table.html
	http://www.chemicalelements.com/groups/halogens.html

OVERVIEW OF THE TOPIC

THE PERIODIC TABLE 9 key areas



PERIODIC	ARRANGEMENT OF	Patterns in the Periodic Table
TRENDS	ELEMENTS IN THE	Electronic structure and the Periodic Table
	PENIODIC PABLE	Properties of transition metals
2 GROUP PROPERTIES	& GROUP!	Properties of Group I elements
	ELEMENTS	Reaction of Group I elements with water
	& GROUP VII	Properties of Group VII elements
	ELEMENTS	Displacement reactions of Group VII elements
		Uses of Group VII elements
	₿ Group 0	Electronic structures of Group 0 elements
	ELEMENTS	Uses of Group 0 elements
	TRENDS	GROUP PROPERTIES GROUP GROUP ELEMENTS GROUP VII ELEMENTS

PERIODIC TRENDS

ARRANGEMENT OF ELEMENTS IN THE PERIODIC TABLE

(a) Patterns in the Periodic Table

- Elements in the Périodic Table are arranged in order of increasing proton number (atomic number).
- A horizontal row in the Periodic Table is known as a period. There are 7 periods in the Periodic Table.
- A vertical column in the Periodic Table is known as a group. There are 8 groups in the Periodic Table.
- The groups in the Periodic table are divided into sections and given names. Fig 1 below shows the sections of the Periodic Table and the names given to them.

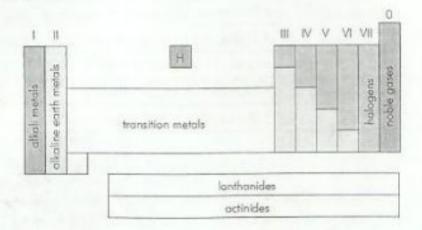


Fig 1: Sections of the Periodic Table

A 'zig-zag' diagonal line in the Periodic table divides metallic elements from non-metallic elements. Elements near the line are called metalloids. Metalloids have the characteristics of both metals and non-metals. Fig 2 shows the position of the diagonal line.

E	11											III		٧	VI	VII	He
Li	Ве						Н					В	C	N	0	F	No
Na	Mg					M	ETAL	S				Al	St	P	S	CI	Ar
K	Ca	Sc	Ti	٧	Cr	Mn	Fe	Со	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
RЬ	Sr	Υ	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	1	Хe
Cs	Ba	Lo	Hf	Ta	W	Re	Os	lr	Pt	Au	Hg	TI	РЬ	81	Po	At	Ro

Fig 2: Diagonal line dividing the metals and non-metals in the Periodic Table

EXAMTED IS

By convention, the group numbers of the elements in the Periodic Table are written in Roman numerals. Metalloids are sometimes called semi-metals.

(b) Electronic Structure and the Periodic Table

Fig 3 below shows the electronic arrangement of the first 20 elements in the Periodic Table.

1	G	iroup Num	han				0
H Herronz prosperient	п.	111	IV	v	VI	VII	4He
7Li Indrove 2.1	9Be	11B Sketning Uniterprised 2.3	12 C filetenec strongerent 2.4	14N Florence strategerent 2.5	160 Federate arrangement 2.6	19 F 1 lectronic consequences 2.7	20 Ne 10 Ne Hactoric protegament 2.8
23Na 11Na thatroni unarganisti 2.8,1	24 Mg 12 Mg (lestrance 7.8.2	27 AI	28 Si 14 Si Christon Christon 2.8.4	31P 15P Electronic unarreparted 2.8.5	32 S 16 S	35.5 CI 17 CI Cleations 2.8.7	40 Ar 18 Ar Flectone 2.8.8
39 K 19 K	40 Ca Distance 2.6.8.2						

Fig 3: Electronic Arrangement of the first 20 elements in the Periodic Table

- Elements in the same period have the same number of electron shells. Going across a period from left to right, the number of outer shell electrons increases by one for every consecutive element ⇒ elements change from metallic to non-metallic character across a period.
- The group number is the same as the number of outer shell electrons in the elements belonging to the same group.
- Elements in the same group have the same number of outer shell electrons (valence electrons). This means that elements in the same group will have similar chemical properties since they will form ions with the same charge. They will also form compounds with similar formulae.
- Going down a group from top to bottom, the number of electron shells increases by one for every consecutive element \Rightarrow elements become more metallic in character, i.e. they lose valence electrons more easily.

EXAMTIP IS

It becomes easier for an element to lose electrons going down a group. With an increase in the number of electron shells, the attraction between the positively charged nucleus and the valence electrons are reduced. The elements become more metallic in character. The element hydrogen is unique because a H atom can form either H', by losing its one valence electron, or H, by gaining one valence electron to complete its outer shell. Forming ions with 1+ charge is characteristic of Group I elements, while forming ions with 1-charge is typical of Group VII elements. This explains why hydrogen is placed by itself in the Periodic Table.

STOP AND THINK

Which one of the following statements concerning the Periodic Table is correct?

- A All groups contain both metals and non-metals,
- B In Group VII, the melting point of the element increases with atomic number.
- C In Group I, reactivity decreases with increasing atomic number.
- D Elements in a period become more metallic with increasing atomic number.
- E Atoms of elements in the same group have the same number of electrons.

EXAMTIP TO

A is wrong because only Groups III, IV, V and VI have both metals and non-metals - the diagonal line cuts through these groups. C is wrong because in Group I, the number of electron shells increase with increasing atomic number. Since the attraction between the nucleus and the valence electron will decrease with more shells, it is easier for a Group I element to lose its valence electron to form stable ions, i.e. reactivity increases down Group I. D is wrong as elements become more non-metallic across a period. Lastly, E is also wrong because atoms of elements in the same group have the same number of valence electrons, not total number of electrons.

[Ans: B] The melting point of Group VII elements increases down the group. Their physical states at room temperature change from gas (fluorine and chlorine) to liquid (bromine)

then solid (iodine).

STOP AND THINK

Phosphorus is in the same group of the Periodic Table as nitrogen. What is the formula of the compound containing only phosphorus and hydrogen?

B PH.

E PH.

EXAM TIP 1 3

Elements in the same group will form compounds with similar chemical formulae. [Ans: C] Nitrogen forms NH, P, being in the same group as N, will form PH,

(c) Properties of Transition Metals

- Transition metals are found in the centre block of Periodic Table.
- They are all hard, strong metals with high melting and boiling points. They also have high density.
- Many transition metals form coloured compounds.
 - Copper(II) compounds are usually blue Examples:

Iron(II) compounds are usually pale green

Iron(III) compounds are usually reddish brown (when solid) or yellow (when in solution)

- They form positively charged ions with variable charges.
 - Copper forms either Cu+ or Cu2+ Examples: Iron forms either Fe2+ or Fe3+
- Many transition metals are used as catalysts in industry to speed up reactions.

The hydrogenation of oil to make margarine uses a nickel catalyst Examples: The manufacture of ammonia uses an iron catalyst

- Many transition metals are also used to make alloys.
 - Steel is made by mixing iron with a small amount of carbon Example:

GROUP PROPERTIES

GROUP I ELEMENTS

(a) Properties of Group I Elements

- Group I elements are also known as alkali metals. They are the most reactive metals in the Periodic Table. These metals are stored under ail or in vacuum to prevent them from reacting with water and/or oxygen in the air.
- Group 1 elements are shiny, silvery sollids.
- Table 1 shows some physical properties of Group I elements.

Name	Symbol	Density (g/cm³)	Melting Point (°C)
Lithium	li	0.53	180
Sodium	Na	0.97	98
Potossium-	K	0.86	64
Rubidium	Rb	1.5	39
Caesium	Cs	1.9	29

Table 1: Physical properties of Group I elements

- Group I metals have low densities and low melting points. They are also soft and easily cut with a knife.
- Going down the group,
 - ① the densities of the elements increase the first 3 elements in the group can float on water.
 - (2) the melting points of the elements decrease.



The compounds of Group I metals are all ionic. Group I metals always form ions with 1+ charge in their compounds.

EXAMTIP TO

Metals in general are hard, dense solids with high melting and boiling points. Group I metals are highly unusual because they are soft, easily cut and have low density and low melting points.

Take note of the trends in the different Groups when studying this topic. Many questions set in the exam involve predicting properties such as melting points and physical states based on the group an element belongs to.

(b) Reaction of Group I Elements with Water

- Group I elements become more reactive down the group.
- Group I elements react with cold water to form metal hydroxides and hydrogen gas.

Group 1 metal + water → metal hydroxide + H₂

2No (s) + $2H_2O$ (l) \rightarrow 2NaOH (aq) + H_2 (g)

The reactivity of Group I metals with water increases down the group:

Li reacts violently

Na reacts very violently, sometimes with an explosion

K reacts explosively

The metal hydroxide solutions formed are all strong alkalis with pH values more than 7.

STOP AND THINK

For the main groups of the Periodic Table, the preperties of the elements show trends down the groups. What is the position of the element which is most metallic in character?

	11	\$ III	IV	V	VI	VII	- 0
6	1	1	1				
1	В		Y	0	D		
A'	1 4	C		Sh.	p-	E	

EXMITTED IN

The most reactive metal is found at the bottom left hand side of the Periodic Table. It is Francium. When it reacts with water, the reaction is very explosive with the production of hydrogen gas and an alkaline solution of francium hydroxide.

[Ans: A] Group I contains the most reactive metals in the Periodic Table.

STOP AND THINK

What describes the solution formed and the gas evolved when potassium reacts with cold water?

	solution	gas
)-9 A	alkaline	alkaline/
- B	akoline	#eutral
C	neutral	alkaline
D	neutral	neutral

EXAMITED IN

The solution formed will turn blue or violet when a few drops of universal indicator is

(A)s: B] Potassium reacts vigorously with water to form potassium hydroxide solution and hydrogen gas.

STOP THINK

Caesium, Cs, is a metal in Group I of the Periodic Table. Which conclusion can be drawn from this information?

A A caesium atom gains an electron when it forms an ion.

B Caesium is more reactive than lithium.

Caesium shows variable valency invits compounds.

Many compounds of caesium are coloured.

EXAMTIP ITS

Transition metals form coloured compounds and show variable valency. The compounds of Group I elements are usually colourless unless the compound contain a transition element. For example, potassium manganate(VII), $KMnO_4$, is purple, and potassium dichromate(VI), K,Cr,O_4 , is prange.

[Ans: B] Group I elements become more reactive down the group. Caesium is expected to be known reactive than lithium.

GROUP VII ELEMENTS

(a) Properties of Group VII Elements

- Group VII elements are also known as halogens. They are very reactive non-metals.
- Table 2 below shows some physical properties of the Group VII elements:

Element	Molecular Formula	Melting Point (°C)	Boiling Point (°C)	Physical State under Room Conditions	Colour
fluorine	F ₂	-220	-148	gas	pale yellow
chlorine	Cl,	-101	-35	gos	yellowish green
bromine	Br ₂	-7	59	liquid	reddish brown
iodine	1,	114	184	solid	block

Table 2: Physical Properties of Group VII elements

- All Group VII elements form diatomic molecules. Each molecule is made up of 2 atoms jained together by a single covalent band.
- Going down the group,
 - the boiling points and melting points increase fluorine and chlorine are gases, bromine is a liquid while iodine is a solid.
 - @ the colour of the elements becomes darker.
- The compounds of Group VII elements can either be ionic or covalent. If the Group VII element is combined with a metal, the compound is ionic. If it is combined with another non-metal, the compound is covalent.

STOP AND THINK

Fluorine, 19 F, is a halogen. The table shows some of the properties of other balogens. Which statement about fluorine is likely to be correct?

element	M,	number of valency electrons	melting point/*C	boiling point/°C	colour of vapour
chlorine	71	7	-101	-35	yellow oreen
bromine	160	7	-7	+59	red-brown
iodine	254	7	+114	+184	violet

- A The boiling point of fluorine is -188°C.
- B The melting point of fluorine is +212°C.
- C The M of fluorine is 19.
- D At r.t.p., fluorine is a black solid

EXAMTIP LO

The element iodine has many colours, depending on what physical state it is in. It is purple in gaseous state, black in solid state, and forms a reddish brown solution when dissolved in water.

[Ans: A] From the table, the boiling point increases from chlorine to iodine mence, the boiling point of fluorine should be lower than that of chlorine.

(b) Displacement Reactions of Group VII Elements

- Group VII elements become less reactive down the group.
 - In a displacement reaction, a more reactive halogen will displace a less reactive halogen from its aqueous salt solution.

Halogen * CI Br The more reactive halogen displaces the less reactive halogen from its aqueous salt solution

Example 1: $Cl_2 + 2Nal \rightarrow 2NaCl + l_2$

Chlorine, being the more reactive halogen, will displace iodine from equeous sodium iodide. The reddish brown solution formed contains sodium chloride and aqueous iodine.

- Example 2: F₂ + 2NaCl → 2NaF + Cl₂
 Fluorine, being more reactive than chlorine, will displace it from aqueous sodium chloride. The yellowish solution formed contains sodium fluoride and chlorine.
- Example 3: There is no reaction between iodine (I₂) and sodium chloride (NaCl). Iodine is less reactive than chlorine and does not displace chlorine from sodium chloride.

EXAMTIP IS

Displacement reactions are redox reactions. In example 1, chlorine is reduced because the exidation state of chlorine is reduced from 0 to -1. Sodium iodide is exidised because the exidation state of iodine increases from -1 to 0.

STOP AND THINK

What deduction about astatine, At, can be made from its position in Group VII?

- A It forms covalent compounds with sodium.
- B It is displaced from aqueous potassium astatide, KAt, by chlorine.
- C It is a gas.
- D It is more reactive than iodine.

EXAMTIP TO

Astatine is the least reactive element in Group VII.

[Ans: B] The reactivity of hologens decreases down the group. More reactive halogens such as chlorine can displace less reactive halogens from a solution of their ions. The equation for this reaction is $Cl_s + 2KAt \rightarrow 2KCl + At_s$

STOP AND THINK

In each of the three experiments, a halogen was added to separate solutions containing ions of one of the other two halogens. The table shows the results. What were the halogens X, Y and Z?

experiment	halogen	halide solution					
	added	X-	γ-	Z-			
1	X,	-	Y, displaced	Z, displaced			
2	Y,	no reaction	- 1	no reaction			
3	Z,	no reaction	Y, displaced	-			

	X	Y	Z
A B C	Br	CI	-
В	Br	- 1	CI
C	CI	Br	-
D	CI	1	Br

EXAMTIP IN

Halogens are axidising agents. When a displacement reaction takes place, the halogen is reduced to its ion - Cl₂ is reduced to Cl-when it takes part in a displacement reaction (see example 1). Since the halogens become less reactive down the group, this implies that their oxidising power decreases as well. Fluorine is thus the strongest axidising agent [Ans. D] X₁ is the most reactive since from experiment 1, it displaces Y₂ and Z₃ from Y and Z₄ respectively.

$$X_2 + 2Y \rightarrow 2X + Y_2$$

 $X_2 + 2Z \rightarrow 2X + Z_1$

Y, is the least reactive since from experiment 2, it cannot displace X_i and Z_i from X and Z' respectively. Hence, X_i must be G_i , and Y_i is I_i .

(c) Uses of Group VII Elements

- Fluorine is used in fluoride toothpaste to help prevent tooth decay.
- Chlorine is used to kill harmful micro-organisms in the water purification process.
- lodine is needed by the human body for proper thyroid gland function.

GROW O ELEMENTS

(a) Electronic Structures of Group 0 Elements

- Group 0 elements are also known as inert gases or noble gases.
- Group 0 elements are the least reactive elements in the Periodic Table because all their outer shells are completely filled. (See topic 3)
- Group 0 elements are monoatomic gases with very low melting and boiling points. They are gases that consist of single atoms.

EXAMTIP 150

The terms 'unreactive' and 'inert' must be distinguished from each other. A substance can be unreactive, but given the correct conditions, it will undergo reaction to form new substances. If a substance is said to be inert, then it is stable and will not take part in reactions no matter what conditions are provided.

(b) Uses of Group 0 Elements

- Argon is used to fill light bulbs to provide an inert atmosphere.
- Neon is used to fill coloured glowing tubes used in advertisements.
- Helium is used to fill weather balloons as it is the second lightest gas.

EXAMTED TOP

Of all the noble gases, argon is the most abundant in air,

SAMPLE QUESTIONS

- 1. Caesium, lithium, potassium and sodium are all in Group I of the Periodic Table.
 - (a) Place these metals in order of reactivity with water, most reactive metal first.
 - (b) Name the chemical products of the reactions between lithium and water and between sodium and water.
 - (c) (i) What would you expect to see if small pieces of caesium were dropped onto water? How would the pH of the resulting solution be different from the pH of water?
 - (ii) Write the full chemical equation for the reaction of caesium with water. Include state symbols.
- 2. Chlorine, bromine and iodine are placed in this order in Group VII of the Periodic Table.
 - (a) State four ways in which the physical or chemical properties of chlorine, bromine and ladine are similar.
 - (b) (i) Describe the trends in physical properties of chlorine, bromine and iodine.
 - (ii) How is the trend in chemical reactivity of chlorine, bromine and iodine shown by displacement reactions? Give an equation for a reaction in which one element displaces another from one of its compounds.
- Fig 4.1 shows a table of elements taken from a larger classification of elements given in the Periodic Table.

- 1	. 11		III	IV	V	VI	VII	0
		hydrogen						helium
lithium	beryllium		boron	carbon	nitrogen	oxygen	fluorine	neon
sodium	magnesium		aluminium	silicon	phosphorus	sulphur	chlorine	argon

- (a) Which group contains
 - (i) halogens
 - (ii) alkali metals?
- (b) From Fig 4.1, choose
 - (i) a metal from Period 2
 - (ii) a non-metal with a valency of 2.
- (c) Hydrogen is difficult to classify into a group as it can be compared with both chlorine and sodium.
 - (i) How many electrons are there in an atom of hydrogen?
 - (ii) What is the valency of a hydrogen atom?
 - (iii) State one way in which hydrogen is similar to chlorine.
 - (iv) State one way in which hydrogen is similar to sadium.
- 4. Oxygen and sulphur are the first two elements in Group VI of the Periodic Table.
 - (a) (i) Name the next element in Group VI.
 - (ii) What is the formula of the compound of this element with hydrogen?
 - (b) At room temperature and pressure, oxygen is a gas and sulphur is a solid.
 - (i) Suggest the state of the element you have named in (a)(i), at room temperature and pressure.
 - (ii) Suggest how the particles in this element are moving at room temperature and pressure.

- 5. The atmosphere on a strange planet consists entirely of a green chemical element. Scientists studying the planet have given the element the chemical symbol Ze. The element does not conduct electricity. It combines with other elements to make covalent and ionic compounds. It forms the ion Ze.
 - (a) From this information give two reason why the element should be classed as a non-metal.
 - (b) Another element occurring on the planet is given the symbol Yn. It has a valency of 3 in its compound with Ze. Write the formula of this compound of Yn and Ze.
 - (c) Suggest
 - (i) the name that Ze is given on Earth
 - (ii) why hydrogen will not burn on the planet to form water.

SOLUTIONS AND EXPLANATIONS

- 1. (a) Caesium, potassium, sodium, lithium
 - (b) When lithium reacts with water, lithium hydroxide and hydrogen are obtained. When sodium reacts with water, sodium hydroxide and hydrogen are obtained.
 - (c) (i) A violent reaction will take place with the production of bubbles of a colourless and odourless gas. A lot of heat and light energy will be produced. The pH of the resulting solution will have a value of around 14 instead of 7, which is the pH of water.
 - (ii) 2Cs (s) + 2H,O (l) → 2CsOH (aq) + H, (g)

EXAMTED TOP .

When the question asks for observations, it is wrong to describe a reaction as "it produces hydrogen gas". This is a deduction obtained by testing the gas evolved with a lighted splint and seeing it extinguish with a 'pop' sound. It is more correct to say that one will 'observe effervescence of a colourless, adounless gas'.

- 2. (a) Similarities in physical properties:
 - (1) All exist as diatomic, covalent molecules.
 - (2) All are as coloured, non-metallic elements.

Similarities in chemical properties:

- (3) All form ionic compounds with metals. Their ions carry 1- charge.
- (4) All form covalent compounds with non-metals.
- (b) (i) Down the group, from chlorine to bromine to iodine,
 - (1) the melting point and bailing point increases
 - (2) the colour intensity of the element increases, i.e. become darker.
 - (ii) In displacement reactions, the more reactive halogen will displace the less reactive halogen from a solution of its ions. An example is chlorine reacting with aqueous potassium lodide solution:

Chlorine, being the more reactive halogen, will displace iodine from aqueous potassium iodide to form aqueous potassium chloride and aqueous iodine.

EXAMTED IN

The way the question is phrased makes it appear very open orded. To score as many marks as possible for such questions without wasting too much time writing excess information, look at the marks allocated to each part. The number of marks allocated will indicate the number of points to include in your answer.

- 3. (a) (i) Group VII
 - (ii) Group I
 - (b) (i) Lithium or beryllium
 - (ii) Oxygen
 - (c) (i) 1
 - (ii) 1
 - (iii) The hydrogen atom can gain one valence electron to form H-, an ion with 1- charge. This is similar to chlorine, which forms CI- during bonding. The CI- ion also has 1- charge.

(iv) Both the hydrogen atom and the sodium atom forms ions with 1+ charge. Hydrogen forms the ion H⁺, while sodium forms the ion Na⁺.

EXAMTIP TO

Other possible similarities between hydrogen and chlorine are: they are both gases, or both have a valency of one. Both sodium and hydrogen also have a valency of one.

- 4. (a) (i) Selenium
 - (ii) H,Se
 - (b) (i) Solid
 - (ii) The particles vibrate in their fixed positions only.

EXAMTIP IN

Since oxygen and sulphur form H₂O and H₂S when they react with hydrogen. Se is expected to form H₂Se, since elements in the same group form compounds with similar formulae.

The melting and boiling points increase down a group. Since sulphur is a solid at room temperature and pressure, selenium must also be a solid as well.

(a) Ze accepts electrons to form Ze⁻. Most non-metals accept electrons to form negatively charged particles.

Ze can form both ionic and covalent compounds. When non-metals combine with metals, they form ionic compounds. When they combine with non-metals, they will form covalent compounds.

- (b) YnZe,
- (c) (i) Chlorine
 - (ii) There is no oxygen on the planet.

EXAMTIP IS

Hydrogen gas must burn in oxygen gas in order to form water. If hydrogen and chlorine were ignited instead, hydrogen chloride will be formed.



METALS

LEARNING OBJECTIVES

Candidates should be able to:

8.1 Properties of Metals

- describe the general physical properties of metals as solids having high melting and boiling points, being malleable and good conductors of heat and electricity, in terms of their structures
- ☑ describe alloys as a mixture of a metal with another element (e.g. brass; stainless steel)
- ☑ identify representations of metals and alloys from diagrams of structures

8.2 Reactivity Series

- place in order of reactivity calcium, copper, (hydrogen), Iron, lead, magnesium, potassium, silver, sodium and zinc, by reference to the reactions, if any, of the metals with water, steam and dilute hydrochloric acid
- Ill deduce the order of reactivity from a given set of experimental results

8.3 Extraction of Metals

- describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series
- describe metal area as a finite resource and hence the need to recycle metals
- discuss the social, economic and environmental advantages and disadvantages of recycling metals

8.4 Iron and Steel

- describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace
- describe steels as alloys which are a mixture of iron with carbon or other metals and how controlled use of these additives changes the properties of the iron (e.g. high carbon steels are strong but brittle, whereas low carbon steels are softer and more easily shaped)
- state the uses of mild steel (e.g. car bodies; machinery), and stainless steel (e.g. chemical plant; cutlery; surgical instruments)

8.5 Prevention of Rusting

- describe the essential conditions for the corrosion (rusting) of iron as the presence of axygen and water; prevention of rusting can be achieved by placing a barrier around the metal (e.g. painting; greasing; plastic coating; galvanising)
- describe the sacrificial protection of iron by a more reactive metal in terms of the reactivity series where the more reactive metals corrodes preferentially (e.g. underwater pipes have a piece of magnesium attached to them)

8.6 Aluminium

state the uses of aluminium and relate the uses to the properties of this metal and its alloys (e.g. the manufacture of aircraft; food containers; electric cables)



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- □ http://www.bbc.co.uk/scotland/education/bitesize/standard/chemistry/metals/index.shtml
- http://www.wpbschoolhouse.btinternet.co.uk/page03/Reactivity.htm
- http://www.s-cool,co.uk/topic_index.asp?subject_id=21
- http://www.letstalkscience.uwo.ca/activities/pdf/chem4.pdf
- □ http://users.argonet.co.uk/users/hoptonj/assets/AS2EX.PDF

OVERVIEW OF THE TOPIC



8.1	PROPERTIES	The Property of the Control of the C	Malleability
	OF METALS	OF METALS	Melting and boiling points
			Conduction of heat and electricity
		& Augys	Formation and structure
			Differences between alloys and pure metals
			Examples of alloys
8.2	REACTIVITY	& ORDER OF REACTIMITY	Reaction of metals with water or steam
	SERIES	@ ORDER OF REACTIVITY	Reaction of metals with dilute hydrochloric acid
			Displacement reactions
8.3	EXTRACTION	& Extraction of Metals	Ease of extraction of metals from their ores
OF METALS		RECYCLING OF METALS	Advantages and disadvantages
8.4	IRON AND STEEL	& EXTRACTION OF IRON	
	SIEEL	Steel	Formation of steels
			High carbon and low carbon steels
			Uses of mild steel and stainless steels
8.5	RUSTING	♣ FORMATION OF RUST	Conditions for rusting
		& PREVENTING THE	Placing a barrier around the iron or steel object
		FORMATION OF RUST	Sacrificial protection
8.6	ALUMINIUM	PROPERTIES OF ALLIMINIUM AND ITS ALLOYS	
		SUSES OF ALUMINIUM AND ITS ALLOYS	

PROPERTIES OF METALS

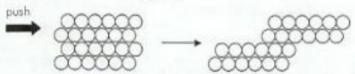
PHYSICAL PROPERTIES OF METALS

Metals usually

- 1 have shiny surfaces -
- are solids
- a have high boiling and melting points
- have high densities
- are good conductors of heat and electricity
- © are strong and tough
- O can be easily bent (malleable) and stretched (ductile) without breaking

Amors

Pure metals are usually too soft and weak for most uses. In pure metals, the atoms are arranged orderly in layers. When a force is applied to the metal, the layers of metal atoms can slide one over another (Fig. 1).



metal structure after a force is applied

Fig 1: The structure of a metal after a force is applied on it

To improve the strength and hardness of pure metals, atoms of another element can be added, usually in small amounts. These atoms prevent the atoms of the metal from sliding over one another, making the metal stronger and harder and less likely to have its shape distorted (Fig 2). The final product is an alloy of the metal.



- one metal, e.g. copper
- second metal, e.g. zinc

Fig 2: The structure of the alloy brass

- An alloy is a mixture of a metal(s) with another element. An example of an alloy is steel which is a mixture of iron and carbon.
- Other advantages of alloying:
 - 1 Improved metal appearance (e.g. pewter = tin + antimony + copper)
 - Increased resistance to corrosion
 (e.g. stainless steel = iron + chromium + nickel + carbon)

STOP AND THINK

An alloy reacted with dilute hydrochlaric acid to evolve a gas which 'pops' a lighted splint. The residue reacted with dilute nitric acid to form a blue solution. Which one of the following pairs of metals is present in the alloy?

- A copper and lead
- C lead and magnesium
- B copper and magnesium
- D lead and zinc
- E magnesium and zinc



Lead will also react with hot dilute hydrochloric acid, but insoluble lead chloride will be formed on the surface of the lead, inhibiting further reaction.

[Ans: B] The blue solution obtained is due to the presence of copper(II) ions, so copper must be present. The other metal present in the alloy reacted with dilute hydrochloric

acid to produce hydrogen gas, so it must be magnesium.

REACTIVITY SERIES



- The reactivity series is a list of metals with the most reactive metal at the top and the least reactive metal at the bottom.
- The order of reactivity can be determined by the reaction of the metal with water or steam and acids. In both types of reaction, if a reaction takes place, hydrogen gas is formed.

(a) Reaction of Metals with Water or Steam

Table 1 lists the reaction of some metals with water.

Metal	Observations/Equation
Potassium	Reacts very violently. Enough heat is produced to ignite the hydrogen gas produced. The hydrogen burns with a blue flame. 2K (s) + 2H ₂ O (l) → 2KOH (aq) + H ₂ (g)
Sodium	Reacts violently. The hydrogen gas produced may catch fire. 2Na (s) + 2H,O (l) → 2NaOH (aq) + H, (g)
Calcium	Reacts readily. Hydrogen gas and calcium hydroxide solution are formed. Ca (s) $+ 2H_2O$ (l) \rightarrow Ca(OH), (aq) $+ H_2$ (g)
Magnesium	Reacts very slowly with cold water. A test-tube of hydrogen gas is produced only after a few days. Mg (s) + 2H _o O (l) → Mg(OH) _o (aq) + H _o (g)
Zinc, Iron, Lead, Copper, Silver	Do not react with cold water.

Table 1: Reaction of metals with water

Table 2 lists the reaction of some metals with steam.

Metal	Observations/Equation		
Magnesium	The hot magnesium reacts violently with steam to form magnesium oxide (a white powder) and hydrogen gas. A bright white glow is produced during the reaction. Mg (s) + H,O (g) → MgO (s) + H, (g)		
Zinc	Hot zinc burns in steam to produce zinc oxide and hydrogen gas. Zinc oxide is yellow when hot and white when cold. $Zn(s) + H_sO(g) \rightarrow ZnO(s) + H_s(g)$		
Iron.	Red hot iron reacts slowly with steam to form hydrogen gas and tri-iron tetraoxide.		
Lead, Copper, Silver	Do not react with steam.		

Table 2: Reaction of metals with steam

When metals react with water or steam, metal hydroxides or metal oxides and hydrogen gas are formed:

metal + water → metal hydroxide + H₂ metal + steam → metal oxide + H₃

- The more vigorous the reaction, the more reactive the metal. Hence, from Tables 1 and 2, we conclude that
 - O Potassium, sodium, calcium are reactive metals.
 - D Magnesium, zinc and iron are fairly reactive metals.
 - Lead, copper and silver are unreactive metals.

EXAMTIP JO

Note that magnesium reacts with both water and steam. When it reacts with water, the product is magnesium hydroxide; when it reacts with steam, the product is magnesium oxide.

(b) Reaction of Metals with Dilute Hydrochloric Acid

Table 3 below lists the reaction of some metals with dilute hydrochloric acid.

Metal	Observations/Equation	
Potassium, sodium Explosive reaction. Reaction is not usually carried out.		
Calcium	Reacts vigorously to give hydrogen gas and calcium chloride. Ca (s) + 2HCl (aq) → CaCl, (aq) + H, (g)	
Magnesium	Reacts rapidly to give hydrogen gas and magnesium chloride, Mg (s) + 2HCl (aq) → MgCl ₂ (aq) + H ₂ (g)	
Zinc	Reacts moderately fast to give hydrogen gas and zinc chloride. Zn (s) + 2HCl (aq) → ZnCl, (aq) + H, (g)	
Iron	Reacts slowly to give hydrogen gas and iron (II) chloride. Fe (s) + 2HCl (aq) → FeCl, (aq) + H, (g)	
Lead, copper, silver	No reaction occurs	

Table 3: Reaction of metals with dilute hydrochloric acid

The list of metals in the reactivity series is shown in Fig 3:

Potassium Sodium Calcium Magnesium	most reactive
Aluminium Zinc Iron Tin	The metal becomes less reactive down the series.
Lead (Hydrogen) Copper Silver Gold	least reactive

Fig 3: The metal reactivity series

Hydrogen is included in the series although it is a non-metal. It serves as a reference point in the series – metals above hydrogen will react with dilute acids to give hydrogen gas, while metals below hydrogen will not react with dilute acids to give produce hydrogen gas.

Aluminium is placed quite high up in the reactivity series although it shows no observable reaction with dilute hydrochloric acid. This apparent lack of reactivity is due to the presence of a layer of aluminium oxide (Al₂O₃) that protects the metal underneath.

EXAM TIP IS

The metal reactivity series may differ from book to book, depending on how many metals are included in it. Generally, Group I metals will be located at the top of the series, since they are the most reactive metals in the Periodic Table. They will be followed by Group III metals, Group III metals and finally, the transition metals.

If a piece of aluminium foil is reacted with hot dilute hydrochloric acid, the initial rate of reaction will be very slow as the ocid reacts with the layer of aluminium oxide on the surface of the fail. Once the oxide layer is removed, the reaction will speed up as aluminium is reactive metal located quite high up in the reactivity series.

STOP AND THINK

Which metal has reacts rapidly with steam, but only very slowly with cold water?

A calcium

C iron

B copper

D potassium

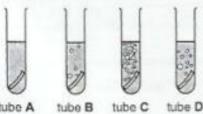


Iron and copper has no reaction with cold water, while potassium will react violently with cold water with vigorous effervescence of hydrogen gas.

[Ans: A] Galcium reacts with cold water to form calcium hydroxide and hydrogen gas, and with steam to form calcium exide and hydrogen gas.

STOP AND THINK

The metals iron, lead, magnesium and zinc are each added to dilute hydrochloric acid. Which tube contains magnesium and dilute hydrochloric acid?



EXAMTIP IS

The metals, arranged in order of decreasing reactivity, are magnesium, zinc, iron and lead. Thus, zinc will be in test tube D, iron in test tube B and lead in test tube A. The reaction between lead and hydrochloric acid is very slow at room temperature. The mixture must be heated in order for reaction to take place.

[Ans: C] Magnesium is the most reactive of all the four metals and should produce the most effervescence.

STOP AND THINK

Which metal has the least tendency to form positive ions?

A calcium

C magnesium

B iron

D sodium

EXAM TIP IS

When sedium undergoes reaction to form a compound, it loses electrons to become sodium ions. The half equation for this change is No -> No - e. Sodium is exidised in the process. All metals undergo exidation when they react to form stable ions.

Metals are also reducing agents. The most powerful reducing agent is found at the top of

the reactivity series.

[Ans: 8] The more reactive metal has the higher the tendency to lose valence electrons and form positive ions. Iron is the least reactive of the four metals and hence forms positive ions least easily.

(c) Displacement Reactions

A more reactive metal will displace a less reactive metal ion from its aqueous solution.

Example 1: Iron + copper(II) nitrate solution \rightarrow iron(II) nitrate solution + copper Fe (s) + Cu(NO₃)₂ (aq) \rightarrow Fe(NO₃)₂ (aq) + Cu (s)

A brown metallic deposit of copper metal will form as the solution turns from blue to pale green due to the formation of iron(II) ions.

Displacement reactions from solutions can be predicted using the reactivity series. A metal higher in the series will displace a metal lower in the series.

Example 2: Iron + zinc(II) sulphate solution
Iron is lower than zinc in the reactivity series. Sin

Iron is lower than zinc in the reactivity series. Since it is less reactive than zinc, no displacement reaction will take place.

Example 3: Copper + silver nitrate solution

Since copper is above silver in the reactivity series, copper will displace silver from silver nitrate solution.

 $Cu(s) + 2AgNO_1(aq) \rightarrow Cu(NO_1), (aq) + 2Ag(s)$

A layer of silver will form on the copper metal. The solution will also turn from colourless to blue due to the formation of copperfill ions.

STOP AND THINK

Strips of metal were dipped into two solutions. A metallic deposit appeared on both strips. What could metal X be?

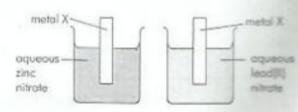
A copper

B iron

C lead

D magnesium

E zinc



EXAMTIP 33

Displacement reactions also take place for Group VII elements. The more reactive halogen will displace the less reactive halogen from a solution containing its ions.

[Ans: D] Metal X must be magnesium because it has to be more reactive than zinc and lead in order to displace them from solutions containing zinc and lead ions respectively.

EXTRACTION OF METALS

EXERTION OF METALS

- [1] There are two methods for extracting metals from their ores:
 - ① Reduction of the metal compound using carbon
 - @ Electrolysis of the molten metal compound
- The method of extraction of a metal from its compounds is determined by its position in the metal reactivity series. The more reactive the metal, the harder it is to extract the metal from its compounds. Fig 4 shows how the extraction method is determined by the reactivity series.

most reactive Reactivity	K Na Ca Mg Al	Extracted by electrolysis of their molten compounds
decreases	Zn Fe Pb Cu	Extracted by reduction of the metal axide using carbon

Fig 4: Extraction of metals and the reactivity series

- Electrolysis involves the use of large amounts of electricity and is a very expensive process compared to reduction using carbon. It is only used to extract very reactive metals because their compounds are too stable to be reduced using carbon.
- When carbon is used to extract a metal from its metal oxide, a redox reaction takes place.

 Carbon is said to be the reducing agent as it reduces the metal oxide to the metal by removing oxygen from it.

Example: copper(II) oxide and carbon

EXAMTER 30

Another reducing agent that is commonly used is hydrogen gas. It is a weaker reducing agent compared to carbon. The reduction of copper(II) exide using hydrogen gas is shown by the equation $CuO + H_s \rightarrow Cu + H_sO$. The hydrogen gas removes exygen from copper(II) exide to form water,

STOP AND THINK

Experiments are carried out to arrange metals X, Y and Z in order of decreasing reactivity. The table shows the results. What is the order of reactivity of the metals?

experiment	X	Y	Z
Does the metal liberate hydrogen from dilute hydrochloric acid?	yes	no	yes
Is the metal oxide reduced by heating with carbon?	yes	yes	no

	most	reactive →	least reactive
Α	Х	Z	Y
В	Y	X	Z
C	Z	X	Y
D	Z	Y	X

EXAMTIP IN

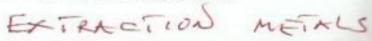
Metals below hydrogen in the reactivity series do not react with dilute oxids to give hydrogen gas. These unreactive metals include copper, silver and gold.

Only metals from zinc downwards in the reactivity series are reduced by heating with carbon. Metals above zinc must be reduced via electrolysis.

[Ans: C] Y is the least reactive metal because it is the only one that does not react with dilute hydrochloric acid. Z is the most reactive metal because it is the only one that is not reduced by heating with carbon.

RECYCLING OF METALS

- Metal are still useful to us.
- The advantages of recycling metals are
 - Better conservation of natural resources, so that reserves last longer. The demand for metal ores will decrease once scrap metal is identified as a viable source of raw material.
 - With recycling, less mining will take place. There will be less air and water pollution caused by the mining process.
 - More effective waste disposal as scrap metal is recovered. Less landfill space will be needed. This will also solve the problem of litter accumulation
- Recycling is sometimes not feasible because of the costs involved. Transportation, sorting through waste and cleaning the scrap metal, etc. may cost more than extracting the metal from its ores. This is true for some cheaper metals.



IRON AND STEEL



BURNON OF BON

- (ii) Iron is extracted from its are haematite by reduction using carbon. The process takes place in a blast furnace (Fig. 5).
- Raw materials:
 - Haematite (containing iron(III) oxide, Fe₂O₃)
 - @ Coke (carbon)
 - 3 Limestone (CaCO₃)

Process:

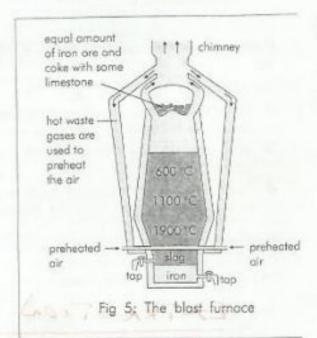
STEP (1) Carbon burns in air to form carbon dioxide:

$$C(s) + O_2(g) \rightarrow CO_3(g)$$

- STEP © Carbon dioxide combines with more coke to form carbon monoxide: $CO_2(g) + C(s) \rightarrow 2CO(g)$
- STEP © Carbon monoxide gas reacts with iron(III) oxide to form moltan iron: $3CO(g) + Fe_2O_3(s) \rightarrow 3CO_2(g) + 2Fe(I)$ Carbon monoxide acts as the reducing agent in the reaction. The liquid iron formed flows to the base of the blast furnace.
- STEP ① Limestone (CaCO₃) decomposes to form calcium oxide and carbon dioxide: $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ The basic calcium oxide is used to remove acidic impurities (e.g. sand, SiO₂). $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$ The liquid slag (CaSiO₃) flows to the base of the blast furnace and floats on top of the molten iron.
- The raw iron obtained in this process is known as cast iron or pig iron. It can be purified further by bubbling oxygen gas through it to burn away impurities. The purified iron is then used to make alloys such as steel.



The reaction between calcium exide and acidic impurites in step 4 is an example of a neutralisation reaction.



STOP AND THINK

Which metal can be extracted by heating an are containing its oxide with carbon?

A calcium

C magnesium

B iron

D potassium

EXAMTIP TOP

Calcium, magnesium and potassium are metals high in the reactivity series. They form very stable compounds that are not reduced by carbon. To extract these metals from their ares, electrolysis is used,

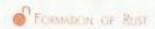
[Ans. 8] Iron is obtained by reducing its ore (haematite) with carbon monoxide gas. The are contains contains iron(III) oxide. The equation for the process is

F ,O + 300 -> 2Fe + 300,

STEELS

- Steel is an alloy made by mixing iron with carbon or other metals. There are many types of steel depending on the type and amount of additives added to it.
- There are 2 kinds of steel: carbon steels and alloy steels.
- Carbon steels contain mainly iron and carbon. They are divided into either low carbon steel (mild steel), or high carbon steel.
 - Description Low carbon steel is softer and more easily shaped. It is used to make car bodies and machinery.
 - ② High carbon steel is strong but brittle. It is used to make knives, hammers, chisels, saws and other cutting and boring tools.
- Alloy steels contain iron and carbon and a transition element such as manganese, nickel, chromium, tungsten or vanadium.
- An example of an alloy steel is stainless steel. It does not rust, is extremely durable and resistant to corrosion even upon heating. It is used to build chemical plants, in making of cutlery and surgical instruments.

RUSTING



- Iron reacts with oxygen in the air in the presence of water to form rust. The chemical formula for rust is hydrated iron(III) axide, Fe₂O₃.xH₂O.
- The overall reaction that takes place is given by the equation

4Fe (s) +
$$3O_2$$
 (g) + $2xH_2O$ (l) \rightarrow $2Fe_2O_3.xH_2O$ (s)

This is an oxidation reaction that takes place slowly. In this process, Iron is first axidised to iron(II) ions before the iron(II) ions are further axidised to iron(III) ions.

$$Fe \rightarrow Fe^{2+} + 2e^{-}$$

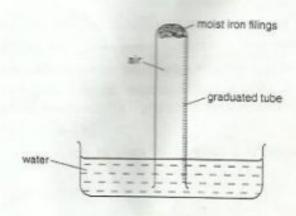
 $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$

Effectively, iron is oxidised into iron(III) ions during rusting.

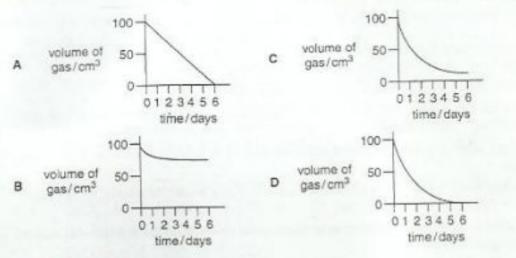
Unlike aluminium which reacts with oxygen in the air to form a protective layer on the metal surface, rust is brittle and flaky. The iron underneath will eventually rust and flake away.

STOP AND THINK

The apparatus shown was set up with 100 cm³ of air in the tube. The volume of gas in the tube was measured at intervals for six days. Which graph best represents how the volume of gas changed over time?







EXAMTIP IP

The rate of rusting will increase in the presence of sodium chloride and acidic pollutants such as SO₂ and CO₂.

[Ans: B] The iron present in steel will rust in the presence of oxygen and water over several days. Since air contains 20% by volume of oxygen, 80% * 100 cm² = 80 cm³ of gas will remain in the test tube after 6 days.

PREVENTING THE FORMATION OF RUST

There are 2 main ways of preventing rusting of iron or steel.

(a) Placing a barrier around the Iron or Steel Object

Coat the iron/steel object with a layer of substance that will stop oxygen in the air and/ or water from reaching the metal

Examples: painting, greasing

(1) Coat the iron/steel object with a less reactive metal or with plastic

Examples: steel food cons coated with tin

(b) Sacrificial Protection

Coat the iron/steel object with a more reactive metal. The more reactive metal will corrode in place of iron.

Example: galvanizing (iron coated with zinc)

- How sacrifical protection works:
 - ① In galvanising, iron metal is plated with zinc metal to prevent the iron from rusting.
 - \oslash When iron corrodes, the first step in the reaction is Fe \rightarrow Fe²⁺ + 2e⁻.
 - 3 Zinc is a more reactive metal than Iron, i.e. zinc loses electrons more easily than Iron.
 - (a) If a piece of zinc is attached to iron, electrons from zinc will flow to the iron metal, preventing it from forming Fe²⁺ ions. The following reaction takes place:

$$Zn \rightarrow Zn^{2+} + 2e^{-}$$

(5) Hence, zinc corrodes in place of iron.

-	
145	GCE 'C' IEVEL
-	Science Chemistry KEY*POINTS

Ship hulls also have bars of zinc attached to it to prevent rusting. Underground pipes are protected using blocks of magnesium connected to them with wires. These blocks of sacrificial metals must be replaced periodically.

EXAMTIP IN

The rate of rusting is decreased when iron is coated or connected via wire to a more reactive metal. If iron is coated with a less reactive metal, as in a tin coated steel food can, the rate of rusting will increase if the coating is scratched. Once the steel is exposed, the iron will start to corrode by forming Fe⁶ ions. Tin, being a less reactive metal than iron, will accept electrons from iron, and this increases the rate of formation of Fe¹ that is found in rust

STOP AND THINK

Which material is not used as a coating to prevent iron and steel from rusting?

A grease

C paint

B magnesium

D zinc



Magnesium is not used as a coating on an iron or steel object because it will react with the oxygen in the air to form magnesium oxide. Magnesium oxide flakes easily and will come off the surface, exposing more magnesium for reaction. In this way, a magnesium coating will wear out very quickly. On the other hand, zinc is coated onto the iron or steel object because the zinc oxide formed when zinc reacts with oxygen in the air will form a coherent layer on the metal and prevent further reaction.

[Ans: B] Magnesium, when used in the sacrificial protection of iron and steel, is usually joined by a wire to the iron or steel object.

8- ALUMINIUM

PROPERTIES OF ALLMINIUM AND ITS ALLOYS

- Aluminjum and its alloys have the following properties:
 - 1 It has low density.
 - D It has good electrical and heat conductivity.
 - (3) It is resistant to corrosion.
 - It is a relatively strong metal.

Uses OF ALUMPHUM AND ITS ALLOYS

Aluminium is used to make food containers, cooking utensils and overhead electrical cables. Its alloy, duralumin, is used to make parts of aircraft bodies and racing cars.

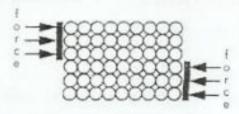
SAMPLE QUESTIONS

- Read the following passage and answer the questions about it.
 - (a) An ore of iron, A, is mixed with coke and limestone and loaded into a blast furnace. Hot B is blasted in through a ring of pipes at the base of the furnace. The coke burns producing gas C which is reduced by reaction with more coke to give gas D. This gas reduces the iron are to iron. The limestone decomposes to firm carbon dioxide and a white solid E which reacts with silica, SiO₂, to give a molten slag containing calcium silicate.
 - (i) Identify A, B, C, D and E
 - (ii) Compound E is basic. What does this tell you about a chemical property of silica?
 - (iii) Write a balanced equation for the reaction between iron(III) oxide and carbon monoxide. Explain why the iron(III) oxide is considered to have been reduced in this reaction.
 - (b) Sodium oxide cannot be reduced to sodium by heating with coke. Instead, sodium is extracted from molten sodium chloride by electrolysis. State why iron can be extracted from its ore in the blast furnace yet sodium cannot.
- 2. Imagine that three metals are transported to Earth by rocket ship from a far distant planet. On the planet the three metals are called beium, ceium and deium. On Earth the three metals are dropped into water: deium does not react but beium and ceium do, liberating a gas which 'pop' when lit. When beium is mixed with dilute sulphuric acid, a solution of beium sulphate is formed. When ceium is dropped into a solution of beium sulphate, beium is not displaced.
 - (a) (i) Name the gas which 'pops' when lit.
 - (ii) Place beium, ceium and deium in order of reactivity, most reactive first.
 - (iii) What would you expect to see if a piece of magnesium was dropped into a solution of deium sulphate?
 - (b) Deium could be the same metal as one of Earth's metals.
 - (i) Name one of Earth's metals that might be the same as deium.
 - (ii) In the future people on Earth might need to import this metal from the far distant planet. Suggest why this may be necessary.
- 3. (a) The way metals are used depends upon their properties. Complete this table.

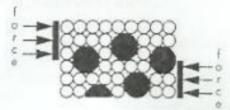
metal	possible use	properties upon which the use depends
aluminium	making parts of aircraft	(i) (ii)
	making saucepans	(iii) (iv)
(v)	galvanising iron	loses electrons more easily than iron

- (b) Name an alloy and the element it contains.
- (c) The circles in the diagrams below represent atoms in a metal when pure and when alloyed. The arrows in the diagrams represent forces applied to the metal and its alloy. Suggest why a force in the directions shown by the arrows will easily change the shape of the pure metal but not the shape of the alloyed metal.

(i) pure metal



(ii) alloyed metal



- (a) Describe how the reactions of the following metals with water may be used to place them in an order of reactivity: calcium, copper, iron and sodium. Write chemical equations for any two of the reactions you describe.
 - (b) The reaction of metal X with water places it between calcium and iron in this order of reactivity. Suggest why X would displace copper if added to a solution of a copper salt.
- This table gives information about six metals. The letters are not the symbols of the metals. Choose the letter of the most suitable metal for each of the uses described below. Give a reason for your choice in each case.

metal	melting point/°C	strength	cost	density/g per cm ¹
Α	1535	strong	inexpensive	7.9
В	3410	strong	expensive	19.4
C	660	strong	fairly expensive	2.7
D	98	weak	inexpensive	0.9
E	961	fairly strong	very expensive	10.5
F	328	fairly strong	inexpensive	11.3

- (a) for the wings and bodies of aircraft
- (b) for the filaments of light bulbs
- (c) for bridges
- (d) for making fishing nets and lines sink

SOLUTIONS AND EXPLANATIONS

1. (a) (i) A is haematite

B is air

C is carbon dioxide

D is carbon monoxide

E is calcium oxide

(ii) Silica is an acidic oxide.

(iii) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Iron(III) oxide has lost oxygen to form iron. It is thus reduced in the reaction.

(b) Sodium is a very reactive metal located high up in the reactivity series. It forms very stable compounds which are not reduced using coke. In comparison, iron is located in the middle of the series and the compounds it forms are not so stable, hence iron can be extracted by reduction using coke.



Always compare the reactivity of metals by referring to their positions in the reactivity series. Simply stating that 'sodium is more reactive than iron' is not good enough.

2. (a) (i) Hydrogen

(ii) beium, ceium, deium

(iii) Magnesium is likely to displace deium metal from the solution. It will dissolve to form magnesium sulphate solution.

(b) (i) Copper

 (ii) Copper metal is a valuable metal and is extensively used. It is a non-renewable resource that is present in very small quantities in copper ores.

EXAMTIP IS

Deium can be any of the metals below hydrogen in the reactivity series.

3. (a) (i) It is light.

(ii) It is strong.

(iii) It is a good conductor of heat.

(iv) It is unreactive

(v) zinc

(b) Brass is an alloy made from copper and zinc.

(c) When a force is applied to the pure metal, layers of metal atoms are able to slide one over another. However, in the alloy, the presence of atoms of different sizes disrupts the regular arrangement of the metal atoms and prevents the layers of metal atoms from sliding.

EXAMTIP IN

Cooking utensils should be made from unreactive metals such as copper so that food cooked using the utensil will not react with it and be contaminated.

No.8 Metals

149 GCE 'O' LEVEL Science (Chemistry) KEY • POINTS

(a) Sadium reacts vigorously with water, giving sodium hydroxide and hydrogen.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Calcium reacts less readily with water. Calcium hydroxide and hydrogen are formed.

Iron does not react with cold water, but will burn in steam to form triiron tetraoxide. Copper does not react with cold water or steam.

The reactivity order is, the most reactive first, sodium, calcium, iron, copper

(b) X is more reactive than iron and iron is more reactive than copper. X is thus more reactive than copper, and will displace copper from a solution of a copper salt.



When iron burns in steam, a reversible reaction takes place where iron combines with steam to give triiron tetraoxide (Fe_1O_2) and hydrogen gas. The same oxide is obtained when iron is burnt in oxygen.

- (a) Metal C because it has a low density. The wings and bodies of aircraft should be strong and light.
 - (b) Metal B because it has a high melting point. Very high temperatures are reached in a filament light bulb and a metal that will remain solid under these conditions is needed.
 - (c) Metal A because it is strong and inexpensive. A bridge must be made of a strong metal that is also inexpensive to reduce cost.
 - (d) Metal F because it has a high density and is inexpensive. Sinks need to be made of a dense metal so that the nets and lines will sink in water, but it cannot be too expensive to be unaffordable.



AIR AND WATER

LEARNING OBJECTIVES

Candidates should be able to:

9.1 Air

- describe the volume composition of gases present in dry air as 79% nitrogen, 20% oxygen and the remainder being noble gases (with argon as the main constituent) and carbon dioxide
- state the uses of oxygen (e.g. in making steel; oxygen tents in hospitals; together with acetylene, in welding)
- name some common atmospheric pollutants (e.g. carbon monoxide, methane, nitrogen oxides (NO and NO₃); ozone; sulphur dioxide; unburnt hydrocarbons)
- state the sources of these pollutants as:
 - (i) carbon monoxide from incomplete combustion of carbon-containing substances
 - (ii) nitragen oxides from lightning activity and internal combustion engines
 - (iii) sulphur dioxide from volcanoes and combustion of fossil fuels
- idiscuss some of the effects of these pollutants on health and on the environment:
 - (i) the poisonous nature of carbon monoxide
 - (ii) the role of nitrogen dioxide and sulphur dioxide in the formation of acid rain and its effects on respiration and buildings

9.2 Water

- outline the purification of the water supply in terms of:
 - (i) filtration to remove solids
 - (ii) chlorination to disinfect the water
- state that seawater can be converted into drinkable water by desalination

USEFUL WEBSITES

- □ http://www.sciencenet.org.uk/database/earth/energy/e00051d.html
 □ http://school.discovery.com/sciencefaircentral/jakesattic/lab/acidrain/p1whatsacidrain.html
 □ http://www.nsc.org/ehc/mobile/airpollu.htm
 □ http://app.nea.gov.sg/cms/htdocs/category_sub.asp?cid=34
- District the state of the state of
- □ http://www.atm.ch.cam.ac.uk/tour/
- □ http://globalwarming.enviroweb.org/
- http://www.helcom.fi/environment/eutrophication.html
- http://www.pub.gov.sg/brought to you.htm

OVERVIEW OF THE TOPIC

AIR AND WATER ** key areas

9.1	AIR	⊕ Сомносттонгог Аія	
		⊕ Oxygen	Determining the percentage of oxygen in air
		+:	Combustion
			Uses of oxygen
		ATMOSPHERIC POUUTION	Common air pollutants
			Sources of air pollutants and their effects of air pollutants on health and on the environment
9.2	WATER	WATER & WATER PURIFICATION	Treatment of raw water
			Desalination

8 AIR

COMPOSITION OF AR

- Clean, dry air is a mixture of gases.
- Table 1 below shows the typical composition of clean, dry air.

Component	Percentage
Nitrogen, N ₂	79%
Oxygen, O ₂	20%
Carbon dioxide, CO ₂	0.04%
Noble gases (mainly argon)	0.93%

Table 1: Composition of clean, dry air

Air also contains water vapour in variable amounts, depending on the humidity of the surroundings.

STOP AND THINK

Air contains about 21% oxygen by volume. When a sample of river water was boiled, the air expelled from the water was found to contain 30%. What is the best explanation of this difference in oxygen content?

- A Carbon dioxide is more soluble in water than is oxygen.
- B Nitrogen reacts with water.
- C Oxygen is more soluble in water than is nitrogen.
- D The noble gases are insoluble in water.

EXAMTIF IS

Air is a mixture of gases and the relative proportions of the component gases are not fixed. The percentages of the gases that make up air will vary slightly from place to place, depending on local conditions.

[Ans: C] Nitregen is less soluble in water than oxygen. This means that more oxygen will dissolve in water, and the relative proportion of oxygen in the air expelled from water is higher.

ONGEN

(a) Determining the Percentage of Oxygen in Air

Fig 1 shows the experimental set-up used to determine the percentage of oxygen by volume in air.

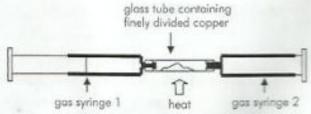


Fig 1: Apparatus for measuring the percentage of oxygen by volume in air

Procedure

- D Gas syringe 1 contains a known amount of air.
- The copper powder is heated strongly as air from gas syringe 1 is passed over it. Copper reacts with oxygen in the air to form copper(II) oxide:

2Cu (s) + O₂ (g) → 2CuO (s)

The remaining gas is collected in gas syringe 2.

- The gas in syringe 2 is passed over the copper powder again to ensure complete reaction.
- When the volume of gas collected in gas syringe 2 remains constant, the apparatus is cooled to room temperature before reading the volume of gas collected.

Volume of
$$O_2$$
 = volume of air in volume of gas collected in syringe 2

%
$$O_2$$
 in air = $\frac{\text{Volume of } O_2}{\text{Volume af air in syringe 1}} \times 100\%$

STOP AND THINK

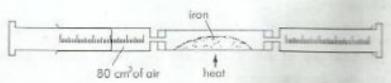
An 80 cm³ sample of air is trapped in a syringe. The air is slowly passed over heated iron in a tube until there is no further decrease in volume. When cooled to the original temperature, what volume of gas remains?

A 16 cm³

B 32 cm³

C 48 cm³

D 64 cm3



EXAMTIP IS

Gases expand and contract on heating and cooling. Hence, the measurement of gas in this experiment must be conducted at the same temperature so that the volumes are occurate. [Ans: D] The remaining gas is mainly nitrogen gas. Since nitrogen gas makes up approximately 80% of air by volume, the volume of nitrogen gas is 80% * 80 cm³ = 64 cm².

(b) Combustion

- Substances react with oxygen to form oxides. This chemical reaction is known as combustion and a lot of heat is generated in the process.
- Complete vs Incomplete combustion of carbon containing substances:
 - When there is a plentiful supply of oxygen, complete combustion occurs and the products are carbon dioxide and water.

Example: Octane, C_8H_{18} (a hydrocarbon found in petrol used in motor vehicles) $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

When there is insufficient supply of oxygen, incomplete combustion takes place. Possible products formed include carbon dioxide, carbon monoxide, water and soot (carbon) particles.

Example:

Octane,
$$C_8H_{18}$$

 $2C_8H_{18} + 19O_2 \rightarrow 4C + 4CO + 8CO_5 + 18H_2O$



Respiration, rusting and combustion are all processes that require axygen.

(c) Uses of Oxygen

- In steel production, oxygen is bubbled into molten iron to burn off impurities.
- Oxygen is used in oxygen tents in hospitals and diving tanks.
- In welding on construction sites, oxygen is mixed with acetylene and burnt to produce a very hot flame used for welding metals together. The flame is called an oxy-acetylene flame.

ATMOSPHERIC POLIUTION

Air is said to be polluted when it contains chemicals in high enough concentrations to harm living things or damage non-living things.

(a) Common Air Pollutants

- Common air pollutants include
 - Carbon monoxide, CO
 - Smoke and dust
 - (3) Oxides of nitrogen (NO and NO₂)
 - (1) Unburnt hydrocarbons, C.H.
 - Sulphur dioxide, 50,
 - Methane, CH,
 - O Ozone, O,

(b) Sources of Air Pollutants and their Effects of Air Pollutants on Health and on the Environment

Table 2 lists the sources of some common air pollutants and their effect on health and on the environment.

Air Pollutant	Sources	Effects
Carbon monoxide, CO	Incomplete combustion of carbon containing substances. Commonly emitted as part of car exhaust fumes.	It is a colourless and adourless paisonous gas. It combines preferentially with haemoglobin in the blood to form carboxyhaemoglobin, which is a very stable compound. This prevents haemoglobin from combining with oxygen, and the body dies from exygen starvation. At lower concentrations, carbon monoxide causes headaches and fatigue.
Nitrogen oxides (nitrogen monoxide, NO, and nitrogen dioxide, NO ₂)	Nitrogen monoxide (NO) is produced when nitrogen (N ₂) reacts with axygen (O ₂) under high temperatures or during lightning storms: N ₂ + O ₂ → 2NO Nitrogen monoxide then undergoes further axidation to form nitrogen diaxide, NO ₂ : 2NO + O ₂ → 2NO ₂ Commonly emitted from car engines and coal burning power stations.	Nitrogen oxides damage lungs and irritate eyes. It also undergoes further reaction with oxygen and rainwater to form acid rain: 4NO₂ + O₂ + 2H₂O → 2HNO₃ Acid rain causes the corrosion of buildings and metals. It also kills plants and marine life by making water bodies become too acidic.

Unburnt hydrocarbons (C _z H _y)	Incomplete combustion of carbon-containing fuels in the internal combustion engines of motor vehicles.	React with other pollutants to form photochemical smag
Sulphur dioxide, SO ₂	Burning of sulphur containing fossil fuels (e.g. coal) and emission during valcano eruption.	Sulphur dioxide causes breathing difficulties and eye initation. It also undergoes further oxidation and dissolves in rain water to form acid rain, 2SO₂ + O₂ + 2H₂O → 2H₂SO₂. Acid rain causes the corrosion of buildings and metals. It also kills plants and marine life by making water badies become too acidic.
Methane, CH ₄	Bacterial decay of vegetable matter, animal dung and rubbish buried in landfills.	It can combine with axides of nitrogen in the presence of sunlight to form photochemical smog. It is also a greenhouse gas that can cause global warming.
Ozone, O ₃	A pollutant formed in a photochemical smag.	At ground level, azone is an eye irritant, causes breathing problems and damages plants

Table 2



Unburnt hydrocarbons, carbon monoxide and nitrogen oxides emitted in car exhaust fumes are removed by the installation of a catalytic converter at the exhaust pipe of the motor car. The catalytic converter uses a platinium catalyst to change these pollutant gases into carbon dioxide, water and nitrogen, which are less harmful to the environment.

STOP AND THINK

A sample of polluted air is bubbled through water. The pH of the solution formed is less than 7. Which gas causes this?

A ammonia

C nitrogen

B carbon monoxide

D sulphur dioxide



Carbon manoxide is also an air pollutant, but it is a neutral exide that is insoluble in water.

[Ans: D] The air pollutant sulphur dioxide is an acidic exide that will dissolve in water to give an acidic solution.

STOP AND THINK

Which substance, present in car exhaust fumes, does not pallute the atmosphere?

A carbon monoxide

C nitrogen oxides

B nitrogen

D unburned hydrocarbans

EXAMTIP IS

Unburnt hydrocarbons and nitrogen oxides combine in the presence of sunlight to produce a phenomenon called photochemical smog. Smogs have been the cause of deaths for the elderly and the young because the heavily polluted air causes breathing difficulties.

[Ans: B] The nitrogen gas is in the air taken in by the engine for fuel combustion. It is emitted as part of the exhaust fumes.

WATER



(a) Treatment of Raw Water

- Purification of drinking water is carried out at the waterworks.
- Three main steps are involved in the water purification process:
 - ⑤ Filtration to remove solids
 - and Addition of activated carbon to remove foul adour and taste.
 - O Chlorination to disinfect the water.
- Fig 2 outlines the water treatment process.

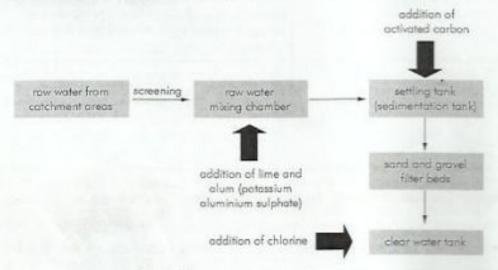


Fig 2: The water treatment process

- STEF © Raw water is first screened to remove large solid impurities. Alum (a coagulating agent) is added to water to cause fine suspended particles to clump together and settle in the sedimentation tank. Lime (calcium oxide) is added to reduce acidity.
- STEF © Activated carbon is added to remove foul adour and taste. Filtration removes any remaining solid particles.
- STEP 3 Chlorination is carried out to kill harmful bacteria present in the water. Fluoride compounds are sometimes also added to help prevent tooth decay.

(b) Desalination

- Desalination is the process of removing dissolved salts from seawater. The sea thus provides a ready source of drinking water.
- Two methods of desalination are commonly used.
 - Distillation: Seawater is evaporated and the pure water vapour formed is condensed, e.g. solar distillation
 - ② Reverse Osmosis: Pure water is extracted from seawater using a semi-permeable membrane under high pressure.



Reverse asmosis is the technology used to obtain Newater

No.9 Air And Water



STOP AND THINK

What is removed in the chlorination stage of the purification of a water supply?

A bacteria

C nitrates

B fluorides

D suspended solids

EXAMTIP IS

In some countries, ozone or UV radiation is used for water sterilization instead.

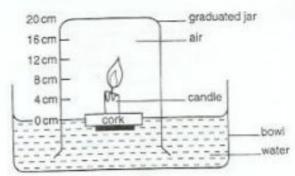
[Ans: A] Chlorine is used to sterilize drinking water by killing the bacteria present in it.

SAMPLE QUESTIONS

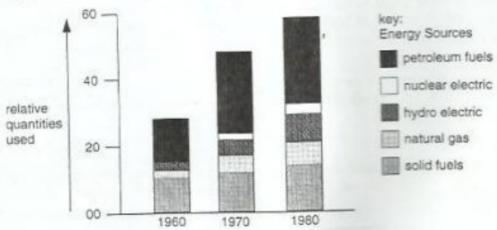
A student wanted to find the percentage of water in air. A wax candle was fixed to a weighted piece
of cork. This was floated on water. The student then lit the candle and lowered a graduated gas
jar over it, as seen in the diagram.

The condle flame slowly went out and the water level rose in the jar. The student noted the final water level in the jar.

- (a) Why did the flame go out?
- (b) What two substances are produced when a wax (hydrocarbon) candle burns?
- (c) What did the water replace when it rose in the jar?
- (d) Why did the water not fill the jar?
- (e) The water rose to the 3.5 cm mark on the jar. Calculate the percentage change of volume of the air.



- (a) Name two goses, other than sulphur dioxide, that pollute the atmosphere. Give one possible source of each pollutant. Sulphur dioxide is shown as an example of a pollutant. pollutant: sulphur dioxide source: burning coal
 - (b) (i) Give one reason for conserving fossil fuels.
 - (ii) What is damaged by the sulphur dioxide released when some fossil fuels are burnt? Give two examples.
- 3. The block graph compares the amounts of different energy sources used by a country in three years.



- (a) What was the main energy source in 1960?
- (b) State one problem of using petroleum fuels at high rates of consumption.
- (c) Which of the sources on the graph is quickly renewed?
- (d) Why does burning large quantities of coal harm plants and animals?

SOLUTIONS AND EXPLANATIONS

- 1. (a) All the oxygen needed to support combustion in the jar is used up.
 - (b) Carbon dioxide and water.
 - (c) The water replaced the oxygen gas used up to burn the candle.
 - (d) Oxygen occupies only about 20% by volume of air. The remaining gas is mainly nitrogen gas.
 - (e) Percentage change = $\frac{3.5}{20} \times 100 = 17.5\%$

EXAMTIP IS

The percentage obtained is less than 20% because the experiment assumes that no product is formed during the combustion. In fact, most of the carbon dioxide formed will dissolve in the water, but some will remain in gaseous state and cause the final water level to be lower than what it should be if no carbon dioxide was present at all.

2. (a) Pollutant: carbon monoxide

Source: incomplete combustion of carbon containing fuels in limited supply of air,

Pollutant: nitrogen dioxide

Source: combustion of nitrogen and axygen under high temperatures in the car engine,

- (b) (i) Fossil fuels are limited, non-renewable energy resources.
 - (ii) Sulphur dioxide in the air will dissolve in rain water to form acid rain. The acid rain will corrode metal and limestone structures. Aquatic plants and animals are also killed when acid rain falls into river bodies because the acid rain makes it too acidic for them to survive.

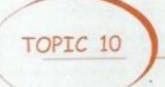


Sulphur dioxide is also an eye irritant and causes breathing difficulties.

- 3. (a) Petroleum fuels
 - (b) It produces carbon dioxide on combustion. Excess carbon dioxide will lead to global warming.
 - (c) Hydroelectric source
 - (d) Coal contains sulphur. Burning large amounts of coal produces sulphur diaxide that will dissolve in rain to form acid rain. When acid rain falls to the ground or into water bodies, the soil and water become too acidic for plants and animals to survive.

EXAMTIP TOP

Excess carbon dioxide causes global warming when it trops too much of the heat that is radiated outwards into space by earth in the night. Global warming may lead to a general rise in ambient temperature, flooding of low lying areas and melting of the polar caps.



ORGANIC CHEMISTRY

LEARNING OBJECTIVES

Candidates should be able to:

10.1 Crude oil

- state that naphtha fraction from crude oil is the main source of hydrocarbons used as the feedstock for the production of a wide range of organic compaunds
- describe the issues relating to the competing uses of oil as an energy source and as a chemical feedstock
- ✓ name natural gas, mainly methane, and petroleum as sources of energy
- describe petroleum as a mixture of hydrocarbons and its separation into useful fractions by fractional distillation
- name the following fractions and state their uses:
 - (i) petrol (gasoline) as a fuel in cars
 - (ii) paraffin (kerosene) as a fuel for heating and cooking and for aircraft engines
 - (iii) diesel as a fuel for diesel engines
 - (iv) lubricating oils as lubricants and as a source of polishes and waxes
 - (v) bitumen for making road surfaces

10.2 Homologous Series

describe a homologous series as a group of compounds with a general formula, similar chemical properties and showing a gradation in physical properties as a result of increase in the size and mass of the malecules (e.g. melting and boiling points; viscosity; flammability)

10.3 Alkanes and Alkenes

- describe the properties of alkanes (exemplified by methane) as being generally unreactive except in terms of burning and substitution by chlorine
- describe the manufacture of alkenes and hydrogen by cracking hydrocarbons and recognise that cracking is essential to match the demand for fractions containing smaller molecules from the refinery process
- describe the difference between saturated and unsaturated hydrocarbons from their molecular structures and by using aqueous bramine
- describe the properties of alkenes in terms of combustion and the addition reactions with bramine, steam and hydrogen
- state the meaning of polyunsaturated when applied to food products
- describe the manufacture of margarine by the addition of hydrogen to unsaturated vegetable oils to form a solid product

10.4 Alcohols

- ☑ describe the properties of alcohols in terms of combustion and oxidation to carboxylic acids
- describe the formation of ethanol by the catalysed addition of steam to ethene and by fermentation of glucose
- state some uses of ethanol (e.g. solvent; fuel; constituent of alcoholic beverages)



10.5 Carboxylic Acids

- describe the formation of ethanoic acid by the oxidation of ethanol by atmospheric axygen or acidified potassium dichromate(VI)
- describe the reaction of ethanoic acid with ethanoi to form the ester, ethyl ethanoate

10.6 Monomers and Polymers

- describe macromolecules as large molecules built up from small units, different macromolecules having different units and/or linkages
- describe the formation of poly(ethene) as an example of addition polymerisation with ethene as the manamer
- state some uses of poly(ethene) as a typical plastic (e.g. plastic bags; cling film)
- describe the structure of the polymer product from a given monomer and vice versa
- describe nylon, a polyamide, and terylene, a polyester, as condensation polymers, the partial structure of nylon being represented as:

and the partial structure of terylene as:

(details of manufacture and mechanisms of these polymerisations are not required)

- state same typical uses of man-made fibres such as nylon and terylene (e.g. clathing; curtain materials, fishing line; parachutes; sleeping bags)
- describe the pollution problems caused by the disposal of non-biodegradable plastics
- identify carbohydrates, proteins and fats as natural macromalecules
- describe proteins as possessing the same amide linkages as nylon, but with different monomer units
- describe fats as esters possessing the same linkages as terylene, but with different monomer units

USEFUL WEBSITES

- □ http://www.schoolscience.co.uk/content/4/chemistry/petrole_m/lowley/index.html
- □ http://www.chemguide.co.uk/organicprops/alkanes/cracking html#top
- ☐ http://www.chemguide.co.uk/organicprops/acids/acidity.htm ≠ top
- ☐ http://www.quia.com/dir/chem/
- □ http://www.rjclarkson.demon.co.uk/junior/organic.htm#Jun%20org%20-%20intro
- ☐ http://www.ausetute.com.au/pblymers.html
- http://www.friendlypackaging.org.uk/schoolplastics.htm

OVERVIEW OF THE TOPIC

ORGANIC CHEMISTRY & key areas

10.1	SOURCES OF	В Інткористіон	A STATE OF THE PROPERTY AND A STATE OF	
	ENERGY	& Petroleum	Fractional distillation	
			Uses of the different petroleum fractions	
10.2	HOMOLOG-	& PROPERTIES OF A	Functional group and general formula	
0.180	OUS SERIES	Homologous Series	Melting and bailing points	
			Viscosity and flammability	
10.3	ALKANES	0.4		
10.0	ALVANCO			
		PROPERTIES OF ALKANES		
		& CHEMICAL REACTIONS OF	Combustion of alkanes	
		ALKANES	Substitution by chlorine	
10.4	ALKENES	& INTRODUCTION		
	P WALL YES			
		PROPERTIES OF ALKENES		
		PREPARATION OF ALKENES	Cracking of large hydrocarban molecules	
		& CHEMICAL REACTIONS OF	Combustion of alkenes	
		ALKENES	Addition of aqueous bromine	
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		TEST FOR UNSATURATION		
10.5	ALCOHOLS	& Introduction		
		& Physical Properties of Alcohols		
		8 PREPARATION OF ETHANOL	Addition of steam to ethene	
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	1	& CHEMICAL REACTIONS OF	Combustion of alcohols	
		ALCOHOLS	Oxidation of alcohals to carboxylic acids	

10.6	CARBOXYLIC ACIDS	& Introduction	Figure 1
	ACIDS	PHYSICAL PROPERTIES OF CARBOXYLIC ACIDS	
		PREPARATION OF ETHANOIC	Oxidation of ethanol in air
		Adb	Oxidation of ethanol using additied potassium dichromate(VI)
		CHEMICAL REACTIONS OF CARBOXYLIC ACIDS	Esterification
10.7	MACRO- MOLECULES	& Introduction	
		SYNTHETIC POLYMERS	Addition polymers
			Condensation polymers
			Uses of synthetic polymers
			 Pollution problems caused by the use of synthetic polymers
		■ NATURAL MACROMOLECULES	Proteins
			• Fats
			Carbohydrates

SOURCES OF ENERGY

INTRODUCTION

- Most of our energy come from the burning of petroleum and natural gas. In some countries, solid coal is used as fuel.
- Petroleum (a dark, sticky liquid also known as crude oil), natural gas and coal are fossil fuels. They are formed as dead plant and animal material are subjected to intense pressure and heat over millions of years. Consequently, these fuels are non-renewable energy sources.
- Natural gas is mostly methane gas (CH₄). It burns cleanly in air to farm carbon dioxide gas and water:

$$CH_{a}(g) + 2O_{2}(g) \rightarrow CO_{2}(g) + 2H_{2}O(g)$$
 $\Delta H = -882kJ$

This reaction is highly exothermic.

Coal is mainly carbon, with small amounts of hydrogen, oxygen, nitrogen and sulphur. When it burns in air, the main products are carbon dioxide and water:

At the same time, small amounts of soot, oxides of sulphur and nitrogen and ash (a solid residue) are formed.

EXAMTED TOP

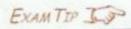
Coal is not a clean fuel. The sulphur dioxide and nitrogen dioxide gases present in the waste gases of a coal burning power station are removed by passing them through wet limestone before the waste gases are emitted into the atmosphere.

PETROLEUM

Petraleum is a mixture of hydrocarbons with different carbon chain length. Malecules with a short carbon chain have low boiling points while those with long carbon chains have high boiling points. Petroleum is quite useless as a mixture; it is usually refined by fractional distillation to separate out its different components to make useful fuels and petrochemicals.

(a) Fractional Distillation

Petroleum is separated into 7 fractions in the fractionating column. Fig 1 shows the different fractions obtained in from a fractionating column.



Students must be able to list the order in which the fractions of petroleum ore obtained.

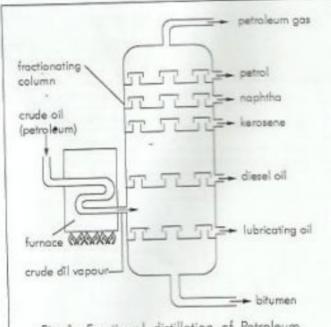


Fig 1: Fractional distillation of Petroleum

(b) Uses of the Different Petroleum Fractions

Table 1 shows the uses for the different petroleum fractions.

	Fraction	Boiling Point (°C)	Length of C chain in Molecule	Uses
top of fractionating column	Petroleum Gas	below toom temperature	1 – 4	bottled gas for gas cooker (liquefied petroleum gas, LPG)
	Petrol (gasoline)	40 - 75	5 – 10	petral for mater vehicles
	Naphtha	75 – 150	7 – 14	chemical feedstock for petrochemical industry, which manufactures plastics, detergents, etc.
	Paraffin (kerosene)	160 – 250	11-16	aircraft engine fuel; fuel for heating and cooking
	Diesel	250 - 300	16 – 20	fuel for diesel engines of heavy vehicles such as lorries, trains, etc.
bottom of	Lubricating Oil	300 – 350	20 – 35	lubricant for moving machine parts; making waxes and polishes
battom of ractionating column	Bitumen	over 350	more than 70 (residue)	road surfacing

Table 1: Uses of petroleum fractions

- About 90% of all the petroleum produced is used as fuel to generate electricity, to drive motor vehicles and to provide power for industries.
- The naphtha fraction is the main source of hydrocarbons used as the chemical feedstock for the production of a wide range of organic compounds.
- To meet the demand for fuel and chemical feedstock, the heavier fractions (fractions with higher boiling points) undergo cracking to provide smaller hydrocarbon molecules, such as those found in petrol.

EXAMTIP IS

To meet the demand for fuel, alternatives such as solar energy nuclear energy and hydroelectric energy can be considered.

STOP AND THINK

When crude oil is distilled, several products are obtained. What is the correct order of their boiling points?

	lowest boiling point		> highes	t boiling point
A	diesel	paraffin	petrol	lubricating oil
В	paraffin	petrol	lubricating oil	diesel
C	petrol	paraffin	diesel	lubricating oil
D	petrol	diesel	lubricating oil	paraffin

EXAMTIP IS

Although crude oil has been separated into 7 fractions, each fraction is not pure. The process is unable to give pure fractions because the boiling points of the hydrocarbons found in crude oil are too close for efficient separation. Hence, in Table 1, the data quoted shows a range of boiling points (indicating the presence of a mixture) instead of a single boiling point.

[Ans: C] When crude oil is fractionally distilled, the fraction with the lowest boiling point is collected from the top of the column, while the fraction with the highest boiling point

is collected at the bottom.

STOP AND THINK

Useful fractions are obtained by the fractional distillation of petroleum. Which fraction and use are correct?

Fraction Use

A bitumen fuel in cars

B lubricating oils for making waxes and polishes

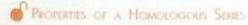
C paraffin (kerosene) for making roads
D petrol (gasoline) aircraft fuel

EXAMTIP TO

When answering questions regarding the uses of the different fractions, do not be vague and give answers like 'kerosene is used in cooking'. Kerosene is inedible, the proper answer should be 'kerosene is used as fuel for cooking'.

[Ans: B] Bitumen is used for making road surfaces, paraffin is used as aircraft fuel while petrol is used as fuel in cars.

HOMOLOGOUS SERIES



- The chemical and physical properties of an organic compound are determined by its functional group. Organic compounds with the same functional group are grouped into a family called a homologous series.
- There are many homologous series and each series is given a name. All organic compounds are named according to the series to which they belong. Table 2 lists some homologous series of organic compounds.
- All homologous series have the following characteristics:
 - 10 They have the same general formula.
 - They have similiar chemical properties because they have the same functional group, i.e. they undergo the same type of reactions.
 - They differ from the next consecutive member by a -CH₂- group.
 - They show a gradual change in physical properties such as melting point, boiling point and density. As the number of carbon atoms in the molecule increases,
 - the melting point, boiling point and density increases.
 - * they become less flammable.

Homologous Series	General Formula	Value of n in General Formula	Functional Group
Alkones	C _n H _{2n+2}	n≥l	ril
Alkenes	C,H _{2n}	n≥2	-C = C -
Akynes	C _n H _{2n-2}	n≥2	-C = C -
Alcohol	C _n H _{2n+1} OH	n ≥ 1	-O-H
Carboxylic acids	C _a H _{2n+1} COOH	n≥0	-C=0 0-H

Table 2: Common homologous series

STOP AND THINK

The diagrams show the structures of four organic molecules. Which two are members of the same homologous series?

D R and S

B P and S

EXAMTIP IN

P and S also differ by a $-CH_2$ - group. Consecutive members of the some hamologous series differ by a $-CH_2$ - group.

[Ans: B] P and S contain the -OH functional group.

Organic compounds are named according to how many carbon they contain and which functional group they possess. Table 3 gives the prefixes assigned to the different number of carbon atoms in a molecule, and Table 4 gives the suffixes assigned to the different homologous series.

Prefix	No. of carbon atoms
meth-	1
eth-	2
pro-	3
but-	4
pent-	5
hex-	6

Table 3: Prefixes in names of organic compounds

- - 5

Suffix	Homologous series	
-one	alkane	
-ene	alkene	
-ol	alcohol	
oic ocid	carboxylic acid	

Table 4: Suffixes in names of organic compounds

- Example 1: An organic molecule belongs to the alcohol series and contains 4 carbon atoms. Since the names of alcohols end with '-ol', the molecule will be called butanol.
- Example 2: The name of the molecule with formula C₂H₃COOH is propanoic acid, since it contains 3 carbon atoms and belongs to the carboxylic acid series.

ALKANES

MINISPUCTION

- The alkanes are a family of hydrocarbons, i.e. they contain hydrogen and carbon atoms only.
- Their general formula is C_nH_{2n+2}, where n is the number of carbon atoms in one molecule.
- Their names end with 'one'.

PROPERTIES OF ALKANES

Table 5 shows the first 5 members of the alkane series.

No. of carbon atoms in one molecule	Chemical name	Molecular formula	Structural formula of one molecule	Relative molecular mass	Physical state at room temperature
1	Methane	CH ₄	H — C — H	16	Gas (–162 °C)
2	Ethane	C ₂ H ₄	H H H H	30	Gas (-89 °C)
3	Propone	C,H _c	$H = \begin{bmatrix} H & H & H \\ I & I & I \\ -C & C & C & -C \\ I & I & I \\ H & H & H \end{bmatrix}$	44	Gas (-42 °C)
4	Butane	C ₄ H ₁₀	H-C-C-C-C-H	58	Gas (-0.5 °C)
5	Pentane	C ₅ H ₁₂	H H H H H H H H H H H H H H H H H H H	72	Liquid (36 °C)

Table 5: Properties of the first 5 members of the alkane series

- Alkanes are covalent compounds with weak intermolecular forces between the molecules.

 As the number of carbon atoms increases, the melting point, and boiling point increase; the first four members are gases, the next thirteen are liquids and the rest are solids.
- Alkanes are insoluble in water.

No.10 Organic Chemistry

169 GCE OF LEVEL SCIENCE (Chemistry KEY-POINTS

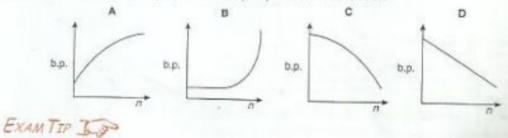
- Alkanes become more viscous, i.e. more difficult to pour out as the number of carbon atoms increase.
- The carbon atoms in alkanes are held together only by single covalent bonds. Alkanes are thus said to be saturated.

EXAMTIP IS

An organic molecule is said to be saturated if it contains only single carbon-carbon covalent bonds. In all organic compounds, each carbon atom will form 4 covalent bonds, while H will form 1 covalent bond. If oxygen atoms are present, each oxygen atom will form 2 covalent bonds.

STOP AND THINK

In the alkane series of hydrocarbons, $C_nH_{2n+2^n}$ the boiling point (b.p.) of the compound increases as n increases. Which graph correctly represents this effect?

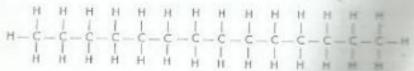


Graph B is wrong because the b.p. is constant for the first few members of the alkane

[Ans: A] As the size of the alkane increases, the strength of the intermolecular bonds between the alkane molecules increases, leading to higher melting and boiling points.

STOP AND THINK

The following displayed formula is of a long chain hydrocarbon. Which set of information is correct about this hydrocarbon?



molecular formula

empirical formula

type of compound

A $C_{14}H_{30}$

CH,

de de compound

B C₁₄H₃₀

C,H,,

okane

 $C = C_{14}H_{30}$ $CH_3(CH_2)_{12}CH_3$

C,H15

polymer

EXAMTER IS

A polymer is a giant molecule made from thousands of atoms joined together by covalent bonds. The alkane $C_{12}H_{10}$ is just a big molecule, not a polymer.

[Ans: B] The molecular formula of the structure is $C_{1}H_{1}$. The simplest ratio of C and H atoms is 7:5, i.e. the empirical formula is $C_{1}H_{1}$. It is an alkane because we can get the formula from the alkane general formula when we scentitute n=7.

CHEMICAL REACTIONS OF ALKANES

(a) Combustion of Alkanes

- Alkanes are not very reactive. They are used mainly as fuels to provide heat energy.
- Alkanes burn in air (oxygen) to form carbon dioxide and water.

Example: methane + oxygen ______ carbon dioxide + water vapour

$$CH_4 + 2O_2 \xrightarrow{burn} CO_2 + 2H_2O$$

When there is not enough air, burning is incomplete. In this case, soot and carbon monoxide are also produced.

Example: ethane + insufficient oxygen ______, carbon + carbon monoxide + carbon diaxide + water vapour

(b) Substitution by Chlorine

Alkanes react with chlorine gas in the presence of sunlight to form chloroalkanes.

Example: methane + chlorine \rightarrow chloromethane + hydrogen chloride CH_a (g) + Cl₂ (g) \rightarrow CH₃Cl (g) + HCl (g)

- This is an example of a substitution reaction. One of the hydrogen atoms in methane has been substituted by a chlorine atom.
- This reaction does not take place in the dark. Sunlight is needed to provide energy to break the CI-CI band to produce chlorine atoms which then react with the alkane molecule.

8- ALKENES



- The alkenes also form a family of hydrocarbons they contain only carbon atoms and hydrogen atoms.
- They have the general formula C_nH_{2n} . Note that the alkene family starts with ethene where n=2. Methene, where n=1 to give the formula CH_2 , does <u>not</u> exist.
- Their names end with 'ene'.
- The formula of each member differs from the previous one by an extra -CH2- group,

PROPERTIES OF ALKENES

Table 6 gives the properties of the first three members of the alkene series.

No. of carbon atoms in one molecule	Chemical name	Molecular formula	Structural formula of one molecule	Relative molecular mass	Physical state at room temperature
2	Ethene	C ₂ H ₄	C-C H	28	Gas
3	Propene	C ₃ H _a	H - C - C - C	42	Gos
4	Butene	C _a H _B	H-C-C-C-C	56	Gas

Table 6: Properties of the first three members of the alkene family

- The alkenes contain carbon-carbon double bonds (-C=C-). This carbon-carbon double bond is known as the functional group of the alkene family. All alkenes must have this functional group.
- Any organic molecule with a carbon-carbon double bond is said to be unsaturated. If a molecule has more than one set of carbon-carbon double bonds, it is said to be poly-unsaturated.

STOP AND THINK

The diagrams show four structures. Which structures represent alkenes with the molecular formula C_4H_8 ?

A 1 and 2 only

C 1, 2, 3 and 4

B 1, 2 and 4 only D 3 and 4 only

EXAMTIP IS

Structures 1, 2, 3 and 4 are called isomers of the compound $C_a H_a$. Isomers are arganic compounds which have the same chemical formula, but different structures. [Ans: 8] Structures 1, 2 and 4 all have the alkers functional group and the correct number of carbon and hydrogen atoms. Although the formula of structure 3 is also $C_a H_a$, it does not have any carbon-carbon double band.

PREMAMBON OF ALIENES

Big hydrocarbon molecules can be broken up into smaller molecules by a process called cracking. The big molecules are passed over a solid catalyst (aluminum oxide or silicon(IV) oxide) at a high temperature (about 600°C), where they break up to give smaller molecules.

When an alkane molecule is cracked, two smaller molecules are produced. One product molecule is an alkane and the other is an alkane.

- Cracking is essential to match the demand for fractions containing smaller molecules from the refinery process. Some of these smaller molecules are used as chemical feedstock, while others are used to produce high grade petrol for motor vehicles.
- Cracking is also used to make hydrogen gas.

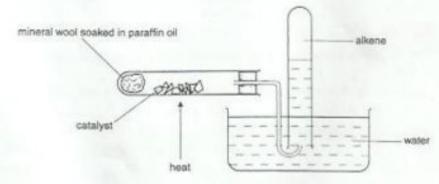
Example:
$$C_{18}H_{38} \xrightarrow{coack} C_8H_{16} + C_{10}H_{20} + H_2$$

The products of cracking cannot be predicted accurately. What we know is that at the end of the process, smaller hydrocarbon molecules (either alkanes or alkenes) and/or hydrogen may be formed.

STOP AND THINK

The apparatus shown is used in the laboratory to form alkenes from paraffin oil. What type of reaction is taking place?

- A combustion
- B cracking
- C distillation
- D reduction



EXAMTIP IS

The catalyst used is often paraus pot. Paraus pot is made from clay, and clay contains silicon dioxide which can be used to speed up the cracking process.

[Ans: B] Paraffin is a heavier fraction containing large hydrocarbon molecules. On heating over a suitable catalyst, it will break down into smaller molecules. These molecules may include alkenes.

CHEMICAL REACTIONS OF ALXENES

Alkenes are more reactive than alkanes. The reactions of alkenes take place at the carboncarbon double bond. During a reaction, the carbon-carbon double bond opens up, allowing the addition of another molecule onto the alkene:

The unsaturation in the alkene molecule is destroyed. A saturated product is formed in which the double bond is replaced by single bonds. An addition reaction is said to have taken

An addition reaction is a reaction in which molecules join together to form a single molecule product.

(a) Combustion of Alkenes

Alkenes burn in plenty of air (oxygen) to form carbon dioxide and water,

ethene + axygen -> carbon diaxide + water vapour $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

Alkenes will produce soot and carbon monoxide when there is insufficient oxygen for complete combustion.

(b) Addition of Aqueous Bromine

Alkenes undergo addition reaction with the aqueous bromine.

Example: C_2H_4 + Br_2 \rightarrow $C_2H_4Br_2$ enhane bronnine 1,2-dibronnoethane

The bromine molecule adds onto the double bond of the ethene molecule.

(c) Addition of Steam

(L) Alkenes react with water (steam) in the presence of phosphoric(V) acid (H₃PO₄) catalyst at high temperature and pressure to form alcohols.

Example:

(d) Addition of Hydrogen

Alkenes undergo addition reaction with hydrogen gas in the presence of a nickel catalyst

Example:

$$C_2H_4$$
 + H_2 \rightarrow C_2H_6 ethane

H H

 I | I

This process is known as hydrogenation. Hydrogenation is used in margarine manufacture to change unsaturated vegetable oils into a solid product. Hydrogen is added onto the double bonds of the vegetable oils.

EXAMTIP ISP

In addition reactions, molecules are always added across a carbon-carbon double bond, i.e. the addition is across adjacent carbon atoms. Hence the final structure of the product will

always take the appearance of $-C=C-+X-Y \rightarrow -C-C-C-$

STOP AND THINK

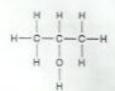
Compound X reacts with steam to produce compound Y. Which structures could represent X and Y?

B H c-c

Another possible structure for Y is

H-C-C-C-H

D



It does not matter which carbon the -OH group attaches itself to, as long as the carbon is one of the carbons from the double bond. The remaining H then attaches itself to the other carbon of the double bond. [Ans: D] X must contain a carbon-carbon double bond to react with steam. Alkenes undergo addition with water to form alcohals.

STOP AND THINK

The diagram shows four stages in a reaction scheme. Which stage involves an addition reaction?



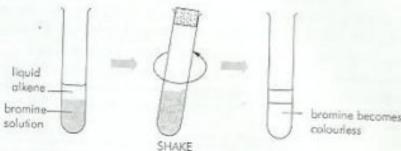
EXAMTIP IN

Stage A involves fractional distillation, stage B involves cracking, and stage D involves oxidation.

[Ans: C] Water must be added to ethene to convert it into ethanol.

TEST FOR UNSATURATION

- We can use the addition reaction as a test to find out if a hydrocarbon is an alkane or an alkane. Fig 2 shows the testing process.
 - (a) liquid alkenes



(b) gaseous alkenes

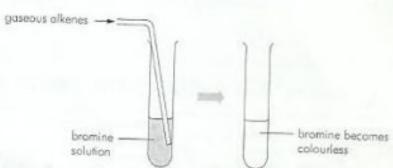
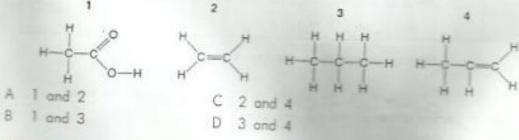


Fig 2 Testing for unsaturation in a molecule

- When an alkene (which contains a C=C bond) is shaken with a solution of bromine in water in a test-tube, the reddish brown colour of bromine is quickly decolourised, i.e. the colour of the mixture in the test-tube changes from reddish brown to colourless. The bromine molecule has added onto the double bond in the alkene molecule.
- When an alkane (which does not have any double band) is shaken with bromine water, there is no reaction. Alkanes do not undergo addition reactions because they are saturated.

STOP AND THINK

The structures of four organic compounds are shown. Which compounds decolourise aqueous



EXAMTED TO

Structure 1 represents a carboxylic acid while structure 3 represents an alkane. Both have

[Ans: C] Both structures 2 and 4 contain carbon-carbon double bonds. When shaken with aqueous bromine, bromine molecules add across the bond and the solution will decolourise.

ALCOHOLS



- The alcohols form a homologous series. All of them contain the hydroxy (-O-H) group of atoms. The -O-H group is also the functional group of the series.
- The general formula for the alcohol series is $C_nH_{2n+1}OH$, where $n \ge 1$. The formula of each alcohol differs from the previous one by an extra $-CH_2-$ group.
- III The names of the alcohols end with 'ol'.

EXAMTIP JUST

Alcohols are not hydrocarbons since they contain oxygen atoms in addition to carbon and hydrogen atoms.

PHYSICAL PROPERTIES OF ALCOHOLS

Table 7 shows the physical properties of the first 4 members of the alcohol series.

No. of carbon atoms in one molecule	Chemical	Molecular formula	Structural formula of one molecule	Relative molecular mass	Physical state at room temperature
1	Methanol	СН₃ОН	H-C-O-H	32	Liquid (bp=64°C)
2	Ethanol	C ₂ H ₆ OH	H-C-C-O-H	46	Liquid (bp=78°C)
3	Propanol	C _s H ₂ OH	H-C-C-O-H	60	Liquid (bp=97°C)
4	Butanol	C ₄ H ₉ OH	H-C-C-C-С-О-Н Н Н Н Н	74	Liquid (op=117°C

Table 7: Physical properties of the first 4 alcohols

- Alcohols are colourless liquids, have low boiling points and are soluble in water.
- As the number of carbon atoms in the alcohol increases,
 - O the boiling point increases
 - The solubility in water decreases
- Alcohols are not alkalis even though they contain the -O-H group. They are neutral liquids.

STOP AND THINK

The formula for butanol may be written as C. H. OH. The next higher member of the homologous series of alcohols is pentanal. How may the formula for pentanal be written?

EXAMTIP IN

Consecutive members of a homologous series differ by a -CH₂- group. Hence, the formula of pentanol can also be worked out by adding a -CH,- group to C,H,OH. [Ans. D] Substituting n = 5 into the general formula for alcohols will give formula D.

STOP AND THINK

Which statement about the homologous series of alcohols is not true?

- A They all contain oxygen.
- B They can be represented by a general formula.
- C They exhibit a gradual change in physical properties.
- D They have the same empirical formula.



Students must be able to differentiate between the terms structural formula, molecular formula, general formula and empirical formula.

[Ans: D] Alcohols do not have the same empirical formula. For example, methanol is CH,OH, or CH,O. Ethanol is C,H,OH or C,H,O. Both formulae cannot be reduced to any simpler



(a) Addition of Steam to Ethene

Ethene is mixed with steam and passed into a reactor containing phosphoric(V) acid catalyst. The reaction takes place at a temperature of 300 °C and a pressure of 65 atm. Addition of water takes place at the double bond and ethanol is formed:

The ethene is obtained by the catalytic cracking of big hydrocarbon molecules that are present in crude oil fractions such as the naphtha fraction.

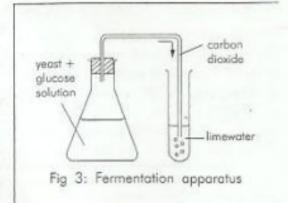
(b) Fermentation of Glucose

Fermentation is a chemical process in which micro-organisms such as yeast act on carbohydrates to produce ethanol and carbon dioxide. Sugars and starch are examples of carbohydrates.

A solution containing glucose (a sugar) is mixed with water and yeast and allowed to react for a few days in the absence of air. During fermentation, the yeast acts on the glucose and converts it into ethanol.

Glucose solution \rightarrow Ethanol + carbon dioxide $C_cH_{12}O_c \rightarrow 2C_2H_5OH + 2CO_2$

The enzymes found in yeast catalyse the reaction. This process takes place at an optimum temperature of 37 °C. If the temperature goes above 40 °C, the enzymes in yeast become denatured so that they can no longer act as catalysts.



- The fermentation of sugars produces only a dilute solution of ethanol (up to 15%). When the ethanol content exceeds this value, the yeast dies and fermentation stops. Higher concentrations of ethanol can be obtained by fractional distillation of the solution.
- The fermentation process must take place in the absence of oxygen. If air is present in the mixture, oxidation of ethanol by the bacteria in the air will take place and the end products will be water and ethanoic acid.

EXAMTED TOP

In the fermentation apparatus above, the limewater has 2 uses. It seeks up the apparatus, preventing air from coming into contact with the reaction mixture. It also indicates the completion of the reaction by turning cloudy due to the carbon dioxide formed during the process.

CHEMICAL REACTIONS OF ALCOHOLS

(a) Combustion of Alcohols

Alcohols burn in plenty of air (oxygen) to give carbon dioxide and water vapour.

Example: Ethanol + oxygen \rightarrow carbon dioxide + water vapour $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

A lot of heat energy is given out in the process. In some countries such as Brazil, ethanol obtained by the fermentation of sugar cane is used as fuel in motor vehicles in place of petrol.

STOP AND THINK

Which equation represents the complete combustion of C3H7OH?

C
$$2C_3H_2OH + 5O_2 \rightarrow 6CO_2 + 8H_2$$

D
$$2C_3H_7OH + 9O_2 \rightarrow 6CO_2 + 8H_2O$$

EXAMTIP JO

In option A, the reaction represented by the equation is an oxidation reaction. Proponal is oxidised into proponal acid (see following section).

[Ans: D] Alcohols burn completely in air to form carbon dioxide and water only.

(b) Oxidation of Alcohols to Carboxylic Acids

Alcohols can be oxidised to carboxylic acids. This reaction takes place in the presence of an oxidising agent such as acidified potassium manganate(VII) [KMnO₄/H⁺] or acidified potassium dichromate(VI) [K₂Cr₂O₃/H⁺].

Example: Ethanol can be oxidised into an organic acid called ethanoic acid.

ethanol +
$$^{'}$$
 2[O] \rightarrow ethanoic acid + $\rm H_2O$

- If the oxidising agent is acidified potassium manganate (VII), the ethanol solution will decolourise it on mixing, i.e. purple turns colourless.
- If the oxidising agent is acidified potassium dichromate (VI), the ethanol will turn it from orange to green on mixing.
- Oxidation of alcohols can also take place they are left exposed to oxygen in the air for a few days. This reaction takes place in the presence of bacteria in the air.

$$C_2H_5OH + 2O_2 \rightarrow CH_3COOH + H_5O$$

This reaction causes wine and beer to sour if they are left exposed to air for sometime. The product is a dilute solution of ethanoic acid called vinegar.

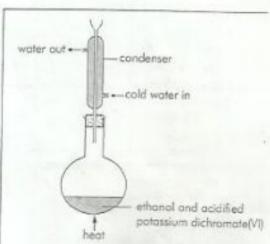


Fig 4: Apparatus set-up for the oxidation of ethanol

EXAMTED TO

In the oxidation apparatus above, the condenser is held vertically to prevent the escape of ethanol vapour. Ethanol has a low boiling point and vaporises easily. When the vapour comes into contact with the cold surface of the condenser, it will liquely and return to the flask.

USES OF ETHANION

- Ethanol is used:

 O as fuel for vehicles
 - as solvent for paints and varnishes
 - 3 in alcoholic drinks such as beer and wine

CARBOXYLIC ACIDS



- The carboxylic acids form a homologous series. All of them have the -C-O-H group of atoms. This group of atoms is the functional group of the series.
- The carboxylic acids have names ending with 'oic acid'. They are so named because these acids can ionise weakly in water to produce H+ ions.
- The general formula for the carboxylic acids is C_nH_{2n+1}COOH, where n = 0 for the first member of the series. The most important carboxylic acid is ethanoic acid.

EXAMTER 300

Since solutions of carboxylic acids are acidic, they will undergo typical reactions of acids - they will react with metals above hydrogen in the reactivity series to form hydrogen, with metal carbonates to form salt, carbon dioxide and water, and with bases to form salt and water.

PHYSICAL PROPERTES OF CARBONNIC ACIDS

Table 8 below shows the physical properties of the first four members of the series.

No. of carbon atoms in one molecule	Chemical name	Molecular formula	Structural formula of one molecule	Relative molecular mass	Boiling Point (°C)
1	Methanoic acid	нсоон	н-с	46	101
2	Ethonoic acid	сн,соон	H-C-C	60	118
3	Propanoic acid	C ₂ H ₅ COOH	H-C-C-C	74	141
4	Butanoic acid	с,н,соон	H-C-C-C-C-C-H	86	164

Table 8: The first 4 members of the carboxylic acid series

The first 4 members are all liquids at room temperature. As the number of carbon atoms in the molecule increase, the boiling point increases.

PREPARATION OF ETHANOIC ACID

Ethanoic acid is prepared by axidation of ethanol.

(a) Oxidation of Ethanol in Air

When a solution of ethanol is exposed to air, the oxygen present slawly axidises ethanol into ethanoic acid in the presence of bacteria. Vinegar, which is a solution of ethanoic acid in water, is made this way.

(b) Oxidation of Ethanol using Acidified Potassium Dichromate(VI)

The orange ocidified potassium dichromate(VI) solution turns green in this reaction.

$$C_2H_3OH + 2[O] \rightarrow CH_3COOH + H_2O$$

EXAMTIP IN

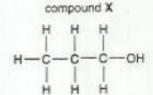
The conversion of ethanol to ethanolic acid is an exidation reaction. Ethanol (C,H,OH) has lost 2 hydrogens and gained 1 exygen in the process.

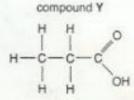
STOP AND THINK

Compound X can be oxidised to compound Y. What are the names of X and Y?

A butanal propanaic acid
B ethanal ethanaic acid

C propanol ethanoic acid
D propanol propanoic acid







When X is exidised to Y, the number of carbon atoms remains unchanged. Only the number of exygen and hydrogen atoms are affected.

[Ans: 51 X is proposed while Y is proposed and Occasio company of the proposed atoms.]

[Ans: D] X is proposed, while Y is proposed acid. Organic compounds containing three carbons have names starting with 'pro-'.

CHEMICAL REACTIONS OF CARBONNIC ACIDS

Carboxylic acids react with alcohols to form esters and water in a reaction called esterification. Concentrated sulphuric acid (H₂SO₄) is used as a catalyst.

Example: ethanoic acid + ethanol ethanoate + water

- In this reaction, ethanoic acid loses the -OH group while ethanol loses the -H group to form water. The remaining sections of the molecules join together to form the ester.
- The name of the ester formed follows the order: alcohol, then acid. For example, if methanol reacts with propanoic acid, the ester formed will be called methyl propanoate.
- Esters are sweet smelling products and are commonly used as food flavourings or as ingredients in perfumes.

EXAMTIP IS

Esterification is not the same as neutralisation even though water is produced in both reactions. In neutralisation, the hydrogen ion reacts with the hydroxide ion to form water. In esterification, an alcohol reacts with a carboxylic acid to form water.

STOP AND THINK

When C₂H₃OH is oxidised it forms substance Z. This substance reacts with C₂H₃OH to produce an ester. What is substance Z?

A HCO,H

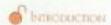
- C CH,CH,CO,H
- B CH,CO,H
- 'D CH,CH,CH,CO,H



Writing formulae the way it is written in the above question is another way of showing the structure of the molecule. If you are unable to recognise the formula, add all the carbons and hydrogens autside the functional group up to help deduce the identity of the compound. C will be C.H.COOH (propancic acid), and D is C.H.COOH (butanoic acid).

[Ans: 8] Ethanol forms ethanoic acid, CH.COOH, when it is exidised.

8-x MACROMOLECULES



Polymers are giant molecules formed by joining smaller molecules together. These smaller molecules are called monomers.

M

-M-M-M-M-M-M-M-Mpolymer molecule

- monomer molecule
- (ii) The process of joining monomers to form a polymer is called polymerisation.
- Polymers contain thousands of atoms linked together by covalent bonds. They are also called macromolecules.
- There are 2 groups of polymers:
 - O Synthetic or man-made polymers (often called plastics), e.g. poly(ethene), terylene and nylon.
 - D Natural polymers found in plants and animals, e.g. proteins, carbohydrates and fats

SYMPHETIC POLYMERS

(a) Addition Polymers

- In addition polymerisation, monomers add onto one another to form a single polymer.
- Example: Formation of poly(ethene)

Polyethene is made from ethene molecules. The molecules contain a carbon carbon double bond (-C=C-) that can add onto one another. The steps below show how to draw the structure of poly(ethene).

STEP ① Draw some ethene molecules side by side:

STEP @ Open the double bonds in the molecules:

STEP @ Join the molecules together.

The structure of poly(ethene) can be represented simply by the structure shown on the right, where n stands for the number of monomers in the structure. The group -CH₂CH₂- in the simplified structure is called the repeating unit of the polymer.

Table 9 shows other examples of addition polymerisation.

Name and formula of monomer	Name and formula of polymer
H CI	H CI
C = C H H	+ C - C + H H H
Chloroethene (Vinyl chloride)	Poly(chloroethene) (PVC or polyvinyl chloride)
F F C C C C C C C C C C C C C C C C C C	
Tetrafluoroethene	Poly(tetraffu oroethene) (PTFE or teflon)
$ \begin{array}{ccc} H & C_0H_5 \\ C &= C \\ H & H \end{array} $	H C ₀ H ₅ -C -C - H H H
Phenylethene (Styrene)	Poly(phenylethene) (Polystyrene)

Table 9: Examples of addition polymerisation

The names of the polymers are obtained by adding 'poly' in front of the name of the monomer in brackets. Hence, if propene undergoes addition polymerisation, the polymer will be called poly(propene).

STOP AND THINK

Which of the following shows the correct monomer for making the polymer?

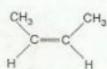
	monomer	polymer
A	H H 	H H H H
-	нн	H H H H
	нн	H H H H
В	H-C-C-O-H	-c-c-o-c-c-o-
С	C = C	H H H H -c-c-c-c-c-
D	CH ₃ H C = C	CH ₃ CH ₃ CH ₃ CH ₃ CH ₃
	CH ₃ H	нини

To join monomers together, they must have either carboncarbon double bonds or reactive functional groups that will link them together on the left as well as on the right to form a chain structure. Structure A does not fulfill the requirements and so will not polymerise. Structure B is only able to link to another molecule on the right but not on the left, so a long chain structure will not be formed. Structure D will give a polymer, but the structure will be

[Ans: C] If the monomers are drawn side by side and the double bands opened up' to link them together, structure B is obtained.

STOP AND THINK

The structure of a polymer is shown. From which hydrocarbon is the polymer made?



D

EXAMTIP IS

To determine the structure of the monomer, first identify the repeating unit in the polymer. The monomer is obtained by closing the ends' of the repeating unit to obtain the double bond. Structure B in the question will be obtained from the repeating unit shown below:

[Ans: B] The polymer is an addition polymer. The monomer from which it is made from must contain a carbon-carbon double bond. From the structure of the repeating unit, the monomer must be B.

(b) Condensation Polymers

- In condensation polymerisation, monomers join together to form a polymer with the elimination of small molecules such as water or ammonia. Terylene and nylon are made using this process.
- Example 1: Formation of terylene

Two different monomers join together to form terylene. The dial and the dicarboxylic acid ends on the monomers can react to form an ester linkage. The steps below show how the polymer is formed.

dio

STEP @ Draw the monomers alternately.

STEP @ Take away one water molecule from each pair of monomer molecules. The molecules are said to condense together to give water.

STEP ① Join the remaining parts of the monomers together.

The repeating unit of terylene is -O-L-0-C-NNV-C-O-Hence, the structure of terylene can be represented as

The units in terylene are joined by the -C-O- group of atoms. We say that terylene has an ester linkage. Polymers containing ester linkages are also known as polyesters.

EXAMTIP TO

The term 'di-' means 'two'. If a molecule is called a dial, it simply means it has two alcohol functional groups. Likewise, a dicarboxylic acid molecule is an acid molecule with two -COOH groups.

Example 2: Formation of Nylon

Two different monomers join together to form mylon. The acid and the amine ends on the monomers can react to form an amide linkage.

STEP ① Draw the monomers alternately.

STEP ② Take away one water molecule from each pair of monomer molecules. The molecules are said to condense together to give water.

STEP 3 Join the remaining parts of the monomers together.

The repeating unit of terylene is
$$-C - \underbrace{N}_{H} - \underbrace{N$$

Hence, the structure of terylene can be represented as

$$\begin{bmatrix} C & C & C & M & M & M \\ C & M & M & M & M \\ M & M & M & M & M \end{bmatrix}$$

The units in terylene are joined by the

We say that nylon has an amide linkage. Polymers containing amide linkages are also known as polyamides.

EXAMTIP IS

Students must be able to identify the ester and amide linkages.

The place where monomers join will also be the place where they will separate if the polymer is broken down into its monomers.

STOP AND THINK

From which pair of reagents could the following polyamide be manufactured? -

A HOOC - COOH and H₂N - NH₂

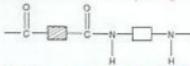
B HOOC -NH, and HOOC -NH,

C HOOC - NH2 and HOOC - NH2

D HOOC COOH and H₂N - NH₂

EXAMTIP TOP

To work out the monomers used to form this polymer, first work out the repeating unit. In this question, the structure of the repeating unit is



This repeating unit contains 2 monomers joined together by an amide band. Break the amide band to get part of the 2 monomers:



During condensation polymerisation, water was removed. Put back the H and OH groups that were removed from the monomers. The OH will go back to the acid (easily identified from the G=O section), and the H will return to the other monomer to give the final structures of the monomers shown in answer D.

[Ans: D] The polymer is an condensation polymer made from two different monomers with diamine and dicarboxylic acid functional groups. (Refer to section on formation of nylon)

(c) Uses of Synthetic Polymers

- All man-made polymers such as poly(ethene), terylene and nylon are called plastics.
- One uses of plastics:
 - Poly(ethene): plastic bags, mineral water bottles, cling film
 - O Nylon: can be made into fibres to make strong ropes (e.g. fishing lines) or woven into cloth to make sleeping bags, parachutes, etc.
 - 3 Terylene: can be made into fibres and woven into cloth

(d) Pollution Problems caused by the Use of Synthetic Polymers

- Plastics burn easily and may produce poisonous gases on combustion. They need to be coated with fire retardants to reduce the risk of fire.
- Plastics are resistant to corrosion. They are also non-biodegradable, i.e. they are not decomposed by bacteria in the ground. Disposal of plastics is difficult and gives rise to environmental pollution when they are incinerated or buried in landfills.

NATURAL MACHOMOLECULES

Proteins, carbohydrates and fots are natural polymers found in plants and animals as well as in our food.

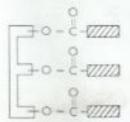
(a) Proteins

- Proteins are needed by both plants and animals mainly for growth, and also to provide enzymes and some energy.
- Proteins are made by polymerising amino acids. The diagram below shows part of the structure of a protein molecule:

Proteins contain the amide linkage, just like nylon.

(b) Fats

- Fats (also vegetable oils) provide the body with energy.
- Fat molecules have the ester linkage, just like terylene.



(c) Carbohydrates

- Carbohydrates provide energy to plants and animals. An example of a carbohydrate is starch.
- Carbohydrates are made from small sugar molecules joined together. The diagram below shows part of the structure of the starch molecule.

STOP AND THINK

Which pair of substances both contain the linkage shown?

- A nylon protein
- B nylon terylene
- C sugars protein
- D terylene poly(ethene)

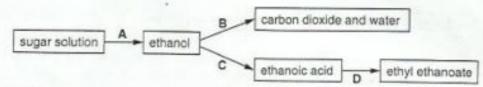
EXAMTED TO

Both polymers are called polyamides. Nylon is a synthetic polyamide, while protein is a natural polyamide.

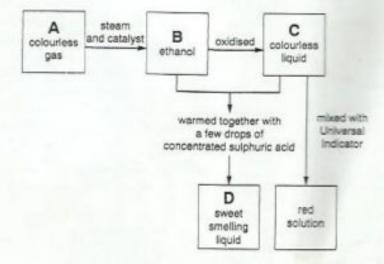
[Ans: A] Both nylon and protein contain the amide linkage.

SAMPLE QUESTIONS

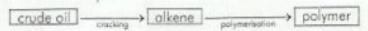
 Chemical reactions sometimes have names. For example, the complete reaction of an acid with an alkali is called 'neutralisation'. The figure contains a series of chemical reactions. Give the names of these reactions.



- (a) reaction A
- (b) reaction B
- (c) reaction C
- (d) reaction D
- 2. (a) Name one homologous series and give the general formula for its members.
 - (b) Explain why, on moving from one member of a homologous series to the next higher member, the relative molecular mass increases by 14. Give an example of another property that changes on moving up the series.
 - (c) The molecules of an alcohol each contain six hydrogen atoms. Suggest the name and formula of this alcohol. Use your knowledge of homologous series to suggest the full structural formula of this alcohol. Determine the percentage of carbon in the alcohol.
- 3. (a) (i) Some hydrocarbons are saturated and others are unsaturated. Name one example of each.
 - (ii) Give one chemical test by which you could distinguish between the two hydrocarbons you have named in (i), and state the result for each substance.
 - (b) Alcohols react with acids to form esters. Ethyl propanoate is an ester.
 - (i) Name two substances that react to form ethyl propanoate.
 - (ii) Give one chemical test by which you could distinguish between the two substances you have named in (i), and state the result for each substance.
- 4. (a) State two uses of ethanol.
 - (b) The diagram below shows some of the properties and reactions of ethanol, B, and some related compounds.



- (i) Identify A, C and D.
- (ii) Write the full structural formula of B, and of any one of A, C or D
- (iii) Write a symbol equation for any one of these reactions.
- 5. This diagram shows two industrial processes.



- (a) What is meant by
 - (i) cracking
 - (ii) polymerisation?
- (b) Give the name and formula of a polymer.
- (c) Some polymers are very unreactive.
 - (i) Suggest a use for a polymer that depends upon this unreactivity.
 - (ii) Suggest a disadvantage of this unreactivity.

SOLUTIONS AND EXPLANATIONS

- (a) fermentation
 - (b) combustion
- (c) oxidation
- (d) esterification

EXAMTIP IS

The equations for the reactions are as follows:

At glucose solution - ethanol + carbon dioxide

B ethanol + oxygen -> carbon dioxide + water

C ethanol - ethanolc acid [oxidation using acidified potassium dichromate(VI)]

D: ethanoic acid + ethanol - ethyl ethanoate + water

- 2. (a) Alkane series. The general formula is C_nH_{2n+2} .
 - (b) Consecutive members of a homologous series differ by a $-CH_2$ group. Relative mass of a $-CH_2$ group = $12 + 2 \times 1 = 14$. The melting and boiling points also increases on moving up a homologous series.
 - (c) Ethanol, C_2H_5OH . The structure of ethanol is H-C-C-O-H

% carbon in ethanol =
$$\frac{2 \times 12}{2 \times 12 + 6 \times 1 + 16} \times 100 = 52.2\%$$

EXAMTIP 1

Other properties that change on moving up a homologous series include viscosity and flammability. Viscosity increases as the number of carbon atoms in the molecule increase, while molecules become more difficult to burn as the number of carbon atoms in them increase.

- (a) (i) saturated hydrocarbon: ethane unsaturated hydrocarbon: ethene
 - Test: Shake both hydrocarbons with aqueous bromine.
 Result with saturated hydrocarbon: No visible reaction. Aqueous bromine remains reddish brown.

Result with unsaturated hydrocarbon: Reddish aqueous bromine decolourised.

(b) (i) alcohol: ethanol

acid: propanoic acid

(ii) Test: Add aqueous sodium carbonate to both substances.

Result for alcohol: No visible reaction.

Result for acid: Effervescence of a colourless, adourless gas that forms white precipitate with limewater observed. Gas is carbon dioxide.



Another possible test to differentiate between the alcohol and the acid is to test using litmus paper. The alcohol is neutral to litmus, but a solution of the acid will turn blue litmus paper red.

- (a) Ethanol is used as a solvent, and as a constituent in alcaholic beverages.
 - (b) (i) A is ethene, C is ethanoic acid and D is ethyl ethanoate.

(iii)
$$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$$

EXAMTIP IS

The structures of A and C are
$$H-C=C-H$$
 and $H-C-C$ respectively. Other equations for (b)(iii) are $CH_1OH+2[O]\to CH_1COOH+H_1O$ and $CH_1OH+CH_1COOH\to GH_1COOC_2H_5+H_2O$

- (a) (i) It is a process whereby large hydrocarbon molecules are broken down into smaller molecules by the use of heat and/or a catalyst.
 - It is a process whereby thousands small molecules are joined together to form a giant molecule.

- (i) The unreactive polymer can be used to make food containers.
 - (ii) The polymer will be non-biodegradable and will contribute to environmental pollution on disposal.

EXAM TIP 3-3

Non biolegradable plastics are often disposed off by burying them in landfills or by burning. Some inert plastics like PVC (poly vinyl chloride) will emit poisonous fumes when burnt, or take up valuable landfill space when they're buried. There is a need to look into recycling of plastics to prevent pollution problems.