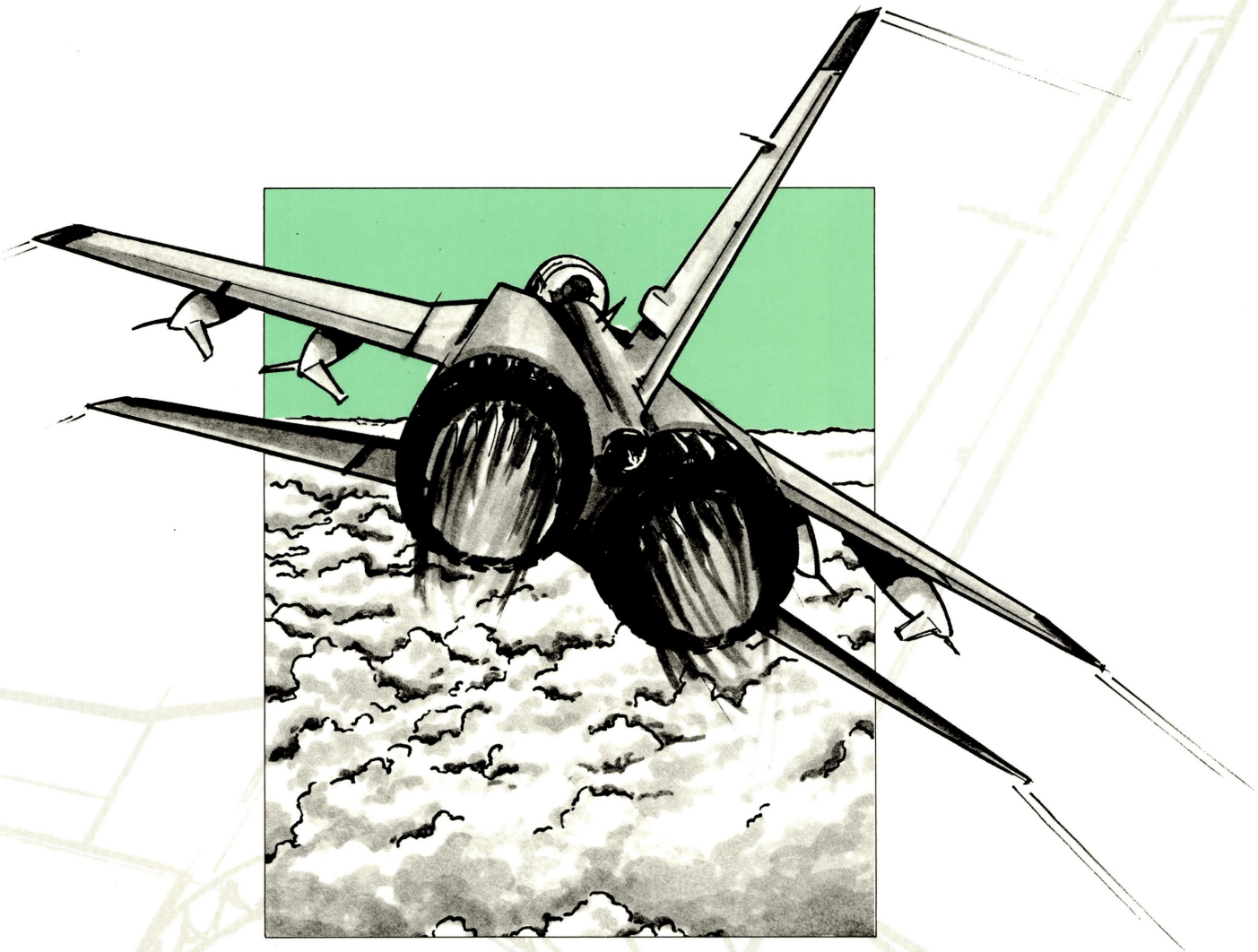
science

THE SALTERS' APPROACH



GCSE UNIT GUIDE

BURNING AND BONDING

Heinemann

NEW EDITION



science

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BURNING AND BONDING

Heinemann

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Introduction

The Salters' Science course for GCSE

The Salters' Science course provides curriculum support materials for a complete programme of balanced science for Key Stage 4.

The Salters' materials may be used as a basis for schemes of work leading to GCSE examinations in Science: Double Award (MEG syllabus 1774), Science: Single Award (1775), Science: Biology (1790), Science: Chemistry (1791) or Science: Physics (1792). This second edition of the course has been extensively revised to match the post-Dearing Programme of Study for Science at KS4, and offers a valuable resource in schools following **any** GCSE science courses.

This book is the guide for teachers for one unit of the GCSE course. It was originally developed by a team of over 30 science teachers working with The University of York Science Education Group. This second edition has been extensively revised by UYSEG, working with a new team who are teaching the course and who have based their work on their experience and on feedback from schools who have been using the course over the past six years.

The activities described in this Unit Guide have been designed to satisfy the following criteria –

- Each topic should have its origins, and hence its justification for study, in aspects of everyday life with which students aged 14–16 years will be familiar either personally or through the media.
- Scientific concepts and explanations should be introduced as they are needed for understanding of the topics being studied.
- Students should be involved in a wide variety of activities.

Industrial, technological, economic, social and environmental aspects of science are, therefore, fundamental to the whole course.

The double award course consists of 23 teaching units. The titles are given on the back cover of this book. Additional units to support the extension material needed for the separate sciences are available from The University of York.

Additional support

Student books, containing one chapter to support each of the course units, are also available. In this Unit Guide, cross-references to the student book material are given at frequent intervals.

GCSE syllabuses, end-of-unit tests for each course unit, sample examination papers and other materials are available from the Midland Examining Group.

INSET support for teachers, including residential courses at York and other centres, is provided at regular intervals.

For further information please contact:

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About this guide

In designing the Unit Guide, the following criteria have been kept in mind –

- New users must be able to acquire quickly an overview of the whole unit and an understanding of the content of each section.
- The guidance must be sufficient to support the inexperienced as well as the experienced teacher. It should be helpful to those who are starting their career and those who are teaching outside their main specialisation.
- The materials must be flexible enough to allow teachers to introduce modifications to, or replacements of, sections or activities where they feel it is appropriate, or to match the needs of students at different levels of attainment.

The next few pages provide general information about this teaching unit.

1 An **overview** which gives a brief summary of the topics covered in the unit.

2 A **syllabus cross-match** chart. For each suggested activity in the unit, cross-references are given to the Key Stage 4 Programme of Study and to assessable outcomes in the GCSE syllabus.

3 A **pre-planner** for the unit. This indicates any requirements which are less usually available or may need advance preparation.

The remainder of the Unit Guide is divided into sections. Each section contains activities which relate to a topic or sub-topic within the unit. A section contains material which may occupy one, two or more lessons.

Each section includes:

4 A **planner** which lists a sequence of suggested teaching and learning activities and identifies the key teaching points covered. It also includes a table of requirements which identifies apparatus and materials required for the activities.

5 **Teachers' Notes** which provide background information, suggestions for managing the activities and for differentiation in the approach and answers to the questions on the Student Sheets.

6 Copy-masters for producing **Student Sheets** which contain information and activities for students.

This Unit Guide is seen as the starting point from which an individual teacher or a department may develop their own scheme of work.

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Overview of the unit

This unit draws together students' ideas about chemical bonding. Students first consider the energy changes associated with bond breaking and bond making. They then revise and extend ideas about types of bonding and the relationships between the type of bonding in a substance, its structure and the resulting properties.

Classification of elements in terms of the types of bonding they show leads to a link between the Periodic Table and the electron arrangements in different elements.

Section 1: *Burning*

Pages 1–18, Student Sheets 1.1–1.8

A study of the Bunsen burner and use of molecular models for hydrocarbon structures leads to a summary of energy changes in bond making and bond breaking. Students carry out experiments which illustrate the idea of activation energy. The flashpoints of different fuels are demonstrated. Students use models to explore the existence of isomers. Students research the products of incomplete combustion and their effects on the atmosphere. A text activity describes the action of three-way catalytic converters in car exhaust systems.

Section 2: *Bonding*

Pages 19–43, Student Sheets 2.1–2.9

This section reviews and extends students' knowledge of the structure of the atom and how atoms bond together. The covalent bond is studied in hydrocarbons, then in allotropes of carbon. An optional activity introduces fullerenes as examples of large covalent molecules. Corrosion of metals provides a context for the study of ionic bonding and a review of the properties of ionic compounds.

Section 3: *A chemical family: the halogens*

Pages 44–65, Student Sheets 3.1–3.6

Students carry out experiments to illustrate the main chemical processes in black and white photography. This introduces the halogens as a family of closely related elements. The oxidising action of the elements is illustrated through their effect on photographic prints. Other properties and uses of the halogens are introduced.

Section 4: *Piecing together the jigsaw*

Pages 66–78, Student Sheets 4.1–4.6

Study of the Periodic Table is extended as students collect information about substances they have studied during the course. Substances are classified according to their structures and these are linked to bulk properties.

The ability of different elements to form ionic and/or covalent bonds is linked to their position in the Periodic Table. Students briefly review the typical properties of elements in Group I and Group II metals, transition metals, the halogens and the noble gases.

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Syllabus cross-match

The table below lists the activities in this unit, section by section. Beside each activity are shown those assessable outcomes from the syllabus content which are addressed through the activity.

The first column identifies the activities. The second column shows the Programme of Study (PoS) section reference which can be used to speed location of the relevant outcomes in the GCSE syllabus booklet. The assessable outcomes are shown in the third column.

Outcomes which are assessed in **both** the Science: Double Award and the Science: Single Award syllabuses are shown in normal type. Outcomes which are assessed **only** in the double award syllabus are shown in *italic type*. Thus, any activity which addresses **only** outcomes shown in *italic type* is **not needed for syllabus coverage for single award** candidates. (However, you may decide to retain some of these activities to provide continuity in the flow of ideas for these students.)

Some syllabus outcomes will only be assessed in the Higher Tier of the examinations. These outcomes are identified by a bullet mark (■) beside the PoS reference code. This indication is given to help when planning schemes of work which include differentiated pathways through the unit. In general, activities which address Higher Tier outcomes may be more suitable for higher-attaining students. However, both the accessibility of an activity and the learning outcomes can be altered by amending the student sheets, or adopting a different approach. Thus, each activity should be considered to see whether it is suitable as written for lower-attaining students, will need to be modified, or might be omitted altogether for such groups.

Particular syllabus outcomes may often be addressed by more than one activity, and sometimes from more than one unit of the course, allowing for revisiting, reinforcement and extension of ideas from one unit to another. The final column shows the letter codes of other units in the course where work on the same or a closely related topic may be found (earlier units where a topic may already have been introduced, or later units where it can be extended or reinforced). These other units should be consulted when developing schemes of work, to decide on the most appropriate level of treatment of the topic in this unit, and how that relates to work done elsewhere in the course.

Section/Activity	PoS	Assessable syllabus outcomes	Unit links
This unit includes a substantial amount of revision of ideas met earlier in the course.	3.1a	know how particles are arranged in solids, liquids and gases	CM A
	3.1b	be able to describe the structure of the atom as a nucleus with external electrons	SB
A number of key topics in chemistry are revisited several times throughout this unit.	*3.1d*	*know the meaning of atomic number, mass number and relative atomic mass for an element*	SB
Syllabus requirements which can be revised repeatedly throughout this unit include:	3.1e ■	be able to construct and interpret diagrams of electron configurations	SB
	3.1l ■	*be able to interpret diagrams showing different types of molecular structures*	CM
	3.2s	know that compounds have a fixed composition and can be represented by formulas	FT TC MM
	3.2s	be able to construct simple symbol equations	
	3.2s ■	*be able to write ionic equations*	MUO

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	Section/Activity	PoS	Assessable syllabus outcomes	Unit links
1	**Practical work 1:** The Bunsen burner	3.2d	understand that the complete combustion of hydrocarbon fuels produces carbon dioxide and water	MUO
		3.2d ■	be able to balance equations for the complete combustion of hydrocarbons	
	Working with models: Breaking and making bonds when methane burns	3.2b	understand the meaning of the term hydrocarbons	MUO
		3.3u	know that chemical reactions are often accompanied by a release of energy	
	Practical work 2: Investigating the energy transfers in reactions	3.3v	understand that some chemical reactions can be monitored by the change in temperature which occurs	
		3.3v	recognise combustion reactions as exothermic processes	
	Practical work 3: Switch off your engine	3.3w ■	be able to draw energy level diagrams for chemical reactions	
		3.3w ■	be able to interpret energy level diagrams in terms of bond breaking and bond making	
		3.3w ■	understand activation energy as the energy needed to start a reaction	
	Demonstration: Flashpoints	3.1a	know how particles are arranged in solids, liquids and gases	
		3.2c ■	be able to explain the boiling range, flammability .... in terms of the strength of forces between molecules	
		3.2c ■	appreciate the importance of flashpoints of liquid fuels in the design of burners and internal combustion engines	
	Text activity 1: Raising the octane number	3.2c	know that natural gas consists mainly of methane and bottled gas is normally propane	MUO
		3.2f	be able to predict the formula of a simple saturated hydrocarbon using the valency of C and H	
		3.2f ■	recognise that methane, ethane and propane are covalently bonded and are members of the alkane series of hydrocarbons	
		3.2e	know that isomers share the same molecular formula but have different molecular structures	
		3.2b ■	know that methane is a major constituent of natural gas, marsh gas and gas from landfill sites where organic matter is decaying	
			know that carbon chains in organic molecules can be straight or branched	
	Group presentations: Pollution from fuels	3.2d	recognise what products may be formed by partial combustion of hydrocarbons	MUO ET
		3.2d	know the gaseous products of burning fossil fuels	A
		3.2w	*understand the main processes by which carbon dioxide is released to the atmosphere*	WW
		3.2d	know that vehicle engines produce carbon dioxide and monoxide, oxides of nitrogen and unburnt hydrocarbons	
	Text activity 2: How do catalytic converters help?	3.2d	appreciate that catalytic converters can reduce unwanted emissions	A WW

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Section/Activity		PoS	Assessable syllabus outcomes	Unit links
1 (cont)		3.2d ■	understand the role of each exhaust gas in contributing to photochemical smog, acid rain and the greenhouse effect	
2	**Text activity 1:** How do atoms bond together?	3.1b	be able to describe the structure of the atom as a nucleus with external electrons	EH SB
		3.1d	*know the meaning of atomic number, mass number and relative atomic mass for an element*	
		3.1e ■	be able to construct and interpret diagrams of electron configurations	
		3.1f	*understand that elements combine to achieve stable outer shells*	
		3.1k ■	*be able to use dot-and-cross diagrams to represent simple covalent structures*	MUO
		3.1k ■	*understand that covalent bonds involve the sharing of electron pairs between atoms*	
	Working with models 1: Different forms of carbon	*3.1l* ■	*know that carbon exists as graphite and diamond which are both giant covalent structures*	
	Text activity 2: Here be diamonds!	*3.1l* ■	*understand the relationship between properties and structure of different forms of carbon*	
	Working with models 2: A new form of carbon to kick around			
	Demonstration: Under attack – clues to another type of bonding	*3.1i*	*know that compounds of metals with non-metals are formed by electron transfer and contain ions*	MM
		3.1i	*understand that metals form positive ions and non-metals form negative ions*	
	Text activity 3: Opposites attract	3.2k	know that metals react with acids (and sometimes with water) to produce hydrogen	CM
		3.2k	know that rusting is an example of an oxidation reaction	
		3.2k ■	know that metals react with oxygen to form basic oxides	
		3.2k ■	*know that metallic oxides are ionic solids*	
		3.1j ■	*understand how ionic lattices can be represented diagrammatically*	
		3.1j	*understand the relationship between structure and properties of ionic compounds*	MM
	Practical work 1: What are ionic compounds like?	*3.1j* ■	*know that electrolysis is the splitting of ionic compounds by electricity*	MM
		3.1j ■	*know that electrolysis can only occur when ions are free to move in melts or in solution*	
	Text activity 4: Some ionic minerals	*3.1i* ■	*understand that current in electrolytes is carried by ions*	MM
	Practical work 2: A different kind of strength			
	Text activity 5: Explaining the special properties of metals	*3.1m* ■	*understand the formation of metallic bonds and the relationship between metallic properties and giant lattices in metals*	

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	Section/Activity	PoS	Assessable syllabus outcomes	Unit links
3	**Discussion:** Images in silver	3.3h	recognise Group VII of the Periodic Table as a family of elements with similar chemical properties	TC
	Practical work 1: Silver halides			
	Practical work 2: Making shadow prints	3.3h	understand the use of silver halides in photography	
	Homework suggestion 1: How are photographic film and paper made?			
	Demonstration: The halogens	3.3h	understand the formation of insoluble salts by precipitation	
	Practical work 3: Now you see it ...			
	Data activity: Close-up on the halogens	3.3h ■	understand the relationship between properties of halogens as non-metallic elements and their electronic configuration	
	Homework suggestion 2: A host of uses			
4	**Discussion 1:** Substances mentioned in this unit	*3.1m* ■	*understand the formation of metallic bonds and the relationship between metallic properties and giant lattices in metals*	
	Classifying data 1: Properties of substances	*3.1m*	*understand the relationship between structure and properties of simple molecules and covalent giant lattices*	CM
		3.1m	*be able to relate properties to structure for substances with linear, layered and cross-linked giant structures*	
	Practical work: Using a key to identify structures	*3.1l* ■	*be able to interpret diagrams showing different types of molecular structures*	CM
	Homework suggestion: Identifying the structures of compounds	*3.1j*	*understand the relationship between structure and properties of ionic compounds*	CM MM
		3.1j ■	*understand how ionic lattices can be represented diagrammatically*	
	Discussion 2: Patterns and the Periodic Table			
	Classifying data 2: Electron arrangements and combining power	3.1f	know how the number of outer electrons for any atom is related to the group in the Periodic Table	TC
		3.1e ■	be able to construct and interpret diagrams of electron configurations	
		3.3b	know that metals have few outer electrons	
		3.3b ■	be able to predict the formula of a compound of an element, given the formulas of compounds of elements in the same group	
		3.3b	know that the combining power of atoms depends on the number of outer electrons	
		3.1f	*understand that elements combine to achieve stable outer shells*	

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Section/Activity		PoS	Assessable syllabus outcomes	Unit links
4 (cont)		3.3a	know that in the Periodic Table: the elements are arranged in order of atomic number; the horizontal rows are called periods; elements with similar properties are in vertical columns called groups; some groups have names – alkali metals, halogens, noble gases	
		3.3j	*know the position of the transition metals in the Periodic Table*	
		3.3j	*know that compounds of transition metals are often coloured*	
		3.3k	*know that transition metals often act as catalysts*	

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Unit pre-planner

Detailed lists of requirements for each activity will be found in the plans for each section of the unit. This sheet merely indicates those requirements which may be less readily available, may take some time to collect or may need preparation in advance.

SECTION	REQUIREMENTS
1	atomic/molecular model kit for construction of models of hydrocarbons samples of petrol and diesel fuels books/leaflets about air pollution
2	models of diamond and graphite structures samples of syntetic diamond (see the additional resources below) junior hacksaw (Eclipse saw) blades piece of galvanised iron
3	camera (not loaded with film) reel of developed film or strip of developed negatives black and white photographs colour photographs efficient blackout facilities medium grade photographic paper commercial photographic developer commercial photographic fixer well-stoppered display samples of chlorine, bromine and iodine a few small iodine crystals in a boiling tube covered with cling film aqueous solutions of chlorine, bromine and iodine (see Teachers' Notes)
4	chemistry textbooks and data books some centres use commercial databases with information about the chemical elements, or set up their own

Additional resources:

Samples of synthetic diamond may be obtained from Molehill Press, Grange Farm, Geddington, Kettering, Northants NN14 1AL.

Leaflets on many aspects of pollution may be obtained from Friends of the Earth, 26–28 Underwood St, London N1 7JQ Tel. 0171 490 1555.

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Burning

In this section students learn about the combustion of hydrocarbons in terms of energy transfers.

2–3 lessons

Some reactions are exothermic and others are endothermic but many require an input of energy (activation energy) to enable them to proceed.

Students also look at the burning properties of petrol, the environmental outcomes of combustion and possible solutions.

ACTIVITY	OUTLINE	KEY POINTS
Practical work 1 (Student Sheet 1.1)	Students investigate gas burner design using Bunsen burners.	Natural gas consists mainly of methane. Hydrocarbon fuels burn to form carbon dioxide and water if sufficient air is present. Incomplete combustion is inefficient and produces products such as soot and carbon monoxide.
Working with models (Student Sheet 1.2)	Students use models to study the breaking and making of bonds during the burning of methane.	A chemical reaction involves bond breaking and bond making.
Practical work 2 (Student Sheets 1.3 and 1.4)	Students investigate energy transfers in chemical reactions.	An exothermic reaction is one which gives out energy for heating the surroundings. An endothermic reaction takes in energy from the surroundings.
Practical work 3 (Student Sheet 1.5)	Students are introduced to the idea that activation energy is needed for a reaction such as combustion to proceed.	Many reactions, including strongly exothermic ones, do not occur until there is an energy input. This is the activation energy.
Demonstration (Student Sheet 1.6)	Students are shown that diesel and petrol have different flashpoints.	The flashpoint is the lowest temperature at which a liquid gives off enough vapour to burn when a flame is applied.
Text activity 1 (Student Sheet 1.7)	The effect of branched-chain alkanes on octane number is used to introduce isomerism.	Carbon chains in molecules can be straight or branched. Isomers have the same molecular formula but different molecular structures.
Group presentations	Students find out the extent to which the use of motor vehicles contributes to the following problems: – Greenhouse effect – Photochemical smog – Acid rain – Health problems Information from different sources is presented to the rest of the class in the form of posters.	Carbon dioxide contributes to the greenhouse effect. Nitrogen oxides, carbon monoxide and unburnt hydrocarbons cause photochemical smog. Nitrogen oxides and sulphur dioxide cause acid rain. Carbon monoxide, nitrogen oxides and lead compounds are toxic and cause health problems.
Text activity 2 (Student Sheet 1.8)	Students learn how some harmful exhaust emissions can be reduced by the use of catalytic converters.	Petrol engines can be fitted with catalytic converters which convert carbon monoxide, unburnt hydrocarbons and nitrogen oxides to carbon dioxide, water and nitrogen.

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ACTIVITY	REQUIREMENTS FOR THE ACTIVITY
Practical work 1	Student Sheet 1.1 *The Bunsen burner* *each group will need:* Bunsen burner, heatproof mat, tripod, gauze, beaker, matches, piece of white porcelain
Working with models	Student Sheet 1.2 *Breaking and making bonds when methane burns* *each group will need:* molecular models of methane and two oxygen molecules
Practical work 2	Student Sheet 1.3 *Investigating the energy transfers in reactions* Student Sheet 1.4 *Results table* **Eye protection must be worn** *each group will need:* watch glass, two test tubes, spatula, thermometer (0–110°C) dropping pipette, 250 cm^3 beaker, safety screens anhydrous copper(II) sulphate (**harmful**), cement powder (**irritant** dust) or plaster of Paris, citric acid, sodium hydrogencarbonate, ammonium nitrate (**oxidising**) *for demonstration only:* 1.0 M hydrochloric acid 1.0 M sodium hydroxide solution (**corrosive**) conc sulphuric acid (**corrosive – always add acid to water**)
Practical work 3	Student Sheet 1.5 *Switch off your engine* *each group will need:* box of matches, Bunsen burner
Demonstration	Student Sheet 1.6 *Flashpoints* *for demonstration:* heatproof mat, long splint, Pyrex test tubes, test tube rack samples of diesel fuel and 40–60 petroleum ether
Text activity 1	Student Sheet 1.7 *Raising the octane number* molecular models (carbon and hydrogen atoms, single bonds)
Group presentations	*each group will need:* access to books/leaflets about air pollution (see the unit pre-planner) poster paper, coloured pens
Text activity 2	Student Sheet 1.8 *How do catalytic converters help?*

Student book: ***Thinking about*** *What happens when a hydrocarbon fuel burns?*



Looking at *Cars and air pollution*

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Burning

Students begin this unit with experiments using the Bunsen burner to illustrate ideas about combustion. This leads to ideas about breaking and making chemical bonds and the associated energy changes, including the concept of activation energy.

The difference between complete and incomplete combustion leads to discussion of some effects of pollutants in the atmosphere.

Practical work 1: *The Bunsen burner*

Student Sheet 1.1 accompanies this activity.

Remind students that natural gas is mainly methane and that in order to use it as a fuel, in a safe and efficient way, a carefully designed burner is needed. The Bunsen burner is a convenient example to investigate.

The students investigate the general principles of gas burner design using Bunsen burners. Since they are asked to dismantle the burners, be sure to supply the type of burner which can be taken apart by unscrewing the chimney from the base. The following points should emerge from their investigations.

- Gas enters the burner through a fine jet.
- Gas mixes with air in the chimney.
- The chimney makes the burner easier to light, and the flame bigger and less likely to blow out.
- The air hole allows variable volumes of air to be mixed with the gas.

The students design experiments to investigate which setting of the air hole gives the cleanest flame, and which the hottest flame. Make sure that what they plan to do is safe before allowing them to proceed. One way of testing the cleanness is to hold a piece of white porcelain, or a piece of glassware, in the flame using tongs. To get a fair comparison, the gas pressure must be kept constant and the flame allowed to play on the object for the same length of time for each flame setting.

One way to compare the heat of the different types of flame would be to measure the rise in temperature of a certain volume of water over a given time. Make sure the students' plans do not involve direct measurement of the flame temperatures using ordinary laboratory thermometers.

This could lead on to a possible investigation using different liquid fuels (usually alcohols). Note that any student plans should be carefully checked for safety.

Answers to questions on *Student Sheet 1.1*

Q1 methane + oxygen → carbon dioxide + water

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

(An experiment to show that these are the products when burning hydrocarbons is done in *Making Use of Oil*, Section 2, Demonstration 2.)

Q2 Carbon (soot), carbon monoxide.

Q3 Increase the supply of air by enlarging the air hole. Decrease the size of the jet to reduce the volume of gas.

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Working with models: *Breaking and making bonds when methane burns*

Student Sheet 1.2 accompanies this activity.

The students use molecular model kits to represent the breaking and making of bonds in the reaction between methane and oxygen. They learn that energy is needed to break bonds in the reactant molecules. In the case of methane and oxygen, more energy is given out when the products are made than was needed to break the bonds in the reactant molecules. It is an example of an exothermic reaction.

The students are introduced to energy level diagrams as a way of representing the energy transfers which occur during a reaction.

Answers to questions on *Student Sheet 1.2*

Q1 2

Q2 2

Q3 2

Q4 More energy is given out than is needed to break bonds in methane and oxygen.

Q5 Overall the bonds in carbon dioxide and water are stronger.

Q6 Any two examples, such as:
alkali metal to water,
acid to water,
thermit reaction, etc.

Q7 Any two examples, such as:
electrochemical cells/batteries,
fireworks, etc.

Practical work 2: *Investigating the energy transfers in reactions*

Student Sheets 1.3 and 1.4 accompany this activity.

Discuss the answers to question 6 on *Student Sheet 1.2*. Other possible examples the students may be familiar with from previous units and KS3 are:

- reactions in our bodies (*Keeping Healthy, Staying Alive, Sports Science*);
- the Haber process for reaction of nitrogen and hydrogen (*Food for Thought*);
- reactions of metals with acids (*Construction Materials, Mining & Minerals*).

Discussion of these or other examples can help to reinforce understanding of the meaning of the term exothermic in terms of bond breaking and bond making. Pose the question: 'Do all reactions give out heat?'

Students then proceed with the practical work given on *Student Sheet 1.3*. A suitable table for recording results is given as *Student Sheet 1.4*.

The first two reactions monitored on *Student Sheet 1.3* are exothermic. The third reaction, citric acid with sodium hydrogencarbonate, is endothermic, dropping in temperature by as much as 12°C.

After discussing the results for these three reactions, demonstrate the following three reactions, getting students to fill in the results of the changes in the table – *Student Sheet 1.4*.

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Demonstration Reaction 5 (students have done reactions 1–4)

This demonstration should be carried out using safety screens and eye protection for staff as well as students.



Note: Always add concentrated sulphuric acid slowly with stirring to a large volume of water when mixing the two (never add water to the concentrated acid).

Put 100 cm^3 of water in a 250 cm^3 beaker and note the temperature. Slowly and carefully add about 2 cm^3 of concentrated sulphuric acid (**corrosive**) to the water. Stir the mixture and measure the temperature again. A temperature rise of 10–12°C should result. At this point, you may wish to emphasise the precautions which are necessary in preparing dilute sulphuric acid from the concentrated acid.

Demonstration Reaction 6

Put 100 cm^3 of water in a 250 cm^3 beaker and note the temperature. Add about 8 g of solid ammonium nitrate (**oxidising**) to the water. Stir the mixture to dissolve the solid and measure the temperature again. A temperature fall of 6–7°C should result.

Demonstration Reaction 7

Put 50 cm^3 of 1.0 M hydrochloric acid in a 250 cm^3 beaker and measure the temperature. Add 50 cm^3 of 1.0 M sodium hydroxide solution (care – **corrosive**), stir the mixture and measure the temperature again. A temperature rise of 6–7°C should result.

Discussion to follow this should include an explanation of endothermic reactions in terms of bond breaking and bond making. An energy diagram for an endothermic reaction may help discussion.

The explanation of energy transfers in chemical reactions in terms of bond breaking and bond making is one that students find very difficult to grasp. This point will have to be continually reinforced throughout the rest of the course.

Representation of the above reactions using energy diagrams similar to the one on Student Sheet 1.5 would help.

Practical work 3: *Switch off your engine*

Student Sheet 1.5 accompanies this activity.

This activity involves the investigation of some exothermic reactions which require an energy input to enable them to proceed.

It is important that the students realise that many reactions do not occur until there is an energy input. The energy level diagram on *Student Sheet 1.5* shows that energy must be supplied for breaking existing bonds before new bonds can begin to form. Students may find it helpful to compare this diagram with the one on *Student Sheet 1.2.*

Students often think that activation energy is simply needed to get a reaction started. In fact, throughout the whole period of the reaction, each molecule requires an energy input to enable it to take part in the reaction. Initially, this energy has to come from an external source (friction lights the match, the match lights the gas). Once many exothermic reactions are started, the energy released by the reaction 'activates' more molecules to keep the reaction going.

Opportunities arise to continually reinforce ideas about bond breaking and bond making.

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Answers to questions on *Student Sheet 1.5*

Q1 Sparks, heat, flare up.

Q2 Given out.

Q3 Less energy.

Q4 Description of activation energy.

Q5 Supply energy to enable the reaction to proceed.

Q6 Description of activation energy. (Involve the idea that bonds have to be broken.)

Q7 The energy given out from the match supplies energy to break bonds in methane and oxygen and so enable the reaction to proceed (activation energy).

Q8 Energy given out is enough to supply activation energy to allow more bonds to be broken, etc.

Demonstration: *Flashpoints*

Student Sheet 1.6 accompanies this activity.

The object of this exercise is to compare the flammability of diesel with that of a petrol substitute and relate the results to the flashpoints of the two fuels. It is illegal to use petrol in schools. 40–60 petroluem ether may be used as a substitute. Petroleum ether has a composition similar to petrol but does not contain benzene. Students may be told either that it is "benzene-free petrol" or "petrol substitute".

Please note that this is a teacher demonstration and should not be done by students.

Wear eye protection. Have safety screen between apparatus and students.
Student must also wear eye protection.
Remove stock bottles from the bench during the demonstration.



Details of the flashpoint test are on *Student Sheet 1.6*.

Make sure, when working with both fuels, that you use small volumes and that there are no naked flames, other than the test flame, nearby. If you follow the prescribed method, petroleum ether vapour will ignite with a faint pop, then the flame will go out. Diesel will not ignite at all.

However, it is important to know how to extinguish the flame in the unlikely event that this becomes necessary; a heatproof mat placed over the mouth of the tube for a few moments is safe and effective. If the vapour ignites when the mat is removed it should be replaced for longer. Do not attempt to extinguish the flame using water.

Flashpoint is defined as the lowest temperature at which a liquid gives off enough vapour to burn when a flame is applied. The flashpoint of diesel is about 100°C, whereas that of petroleum ether is about –40°C. Therefore, when the flashpoint test is carried out at room temperature, petroleum ether ignites but diesel does not.

Explain to the students that this is one factor which has led to differences in the design of petrol and diesel engines.

The students are asked to answer questions about the nature of the intermolecular forces present in diesel and petrol.

Answers to questions on *Student Sheet 1.6*

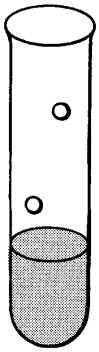
Q1 The flashpoint of diesel fuel is higher than room temperature.

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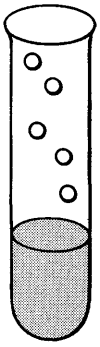
Q2 Petrol's flashpoint is lower than that of diesel, and below room temperature.

Diagrams representing the behaviour of diesel fuel can be drawn as follows:

(a) at room temperature



(b) at diesel's flashpoint



Q3 The diesel molecules are held more strongly in liquid diesel than the petrol molecules in liquid petrol.

Q4 Petrol boils at a lower temperature than diesel. This means that at a given temperature there is more vapour above the liquid, so the vapour pressure is higher. A liquid boils when its vapour pressure equals the external pressure; petrol will achieve this vapour pressure at a lower temperature than diesel.

Text activity 1: *Raising the octane number*

Student Sheet 1.7 accompanies this activity.

Students are introduced to structural isomerism by learning about how the octane number of petrol can be increased by adding more branched-chain alkanes.

Molecular models have been used before in this section and were used to represent hydrocarbons in *Making Use of Oil*. Students should be encouraged to make models of all their proposed isomers and compare their shapes.

Groups work step by step through methane, ethane and propane, then explore the possibility of making two different arrangements for C_4H_{10}, and three different structures for C_5H_{12}. This will allow them to become familiar with the models, and the valency rules, before their attention turns to the problem of comparing different structures for isomers.

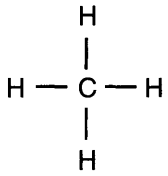
It is wise to check that students have remembered the rules of four links for each carbon atom and one for each hydrogen. Faster working groups might be allowed to go on to explore the possibilities for C_6H_{14}.

When students are writing down structural formulas to represent isomers, they very easily become confused. Representations showing a methyl side group on the second or third carbon along a four-carbon chain look like different structures. However, if students can see that turning a model end-over-end converts one layout to the other, they may find it easier to work out the right combinations.

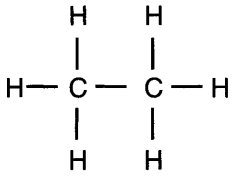
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Answers to questions on *Student Sheet 1.7*

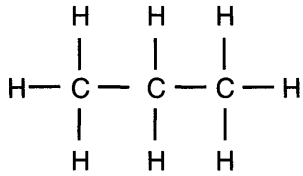
The model structures should be as follows:



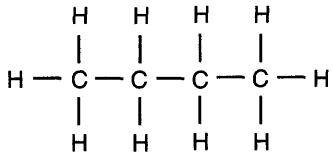
methane



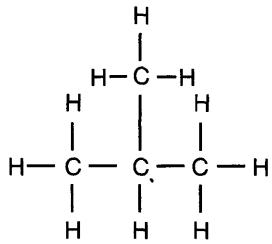
ethane



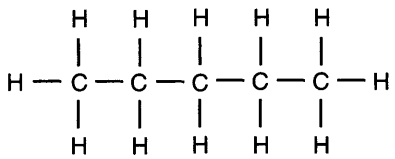
propane



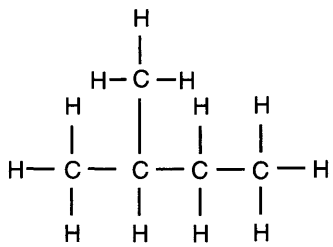
butane



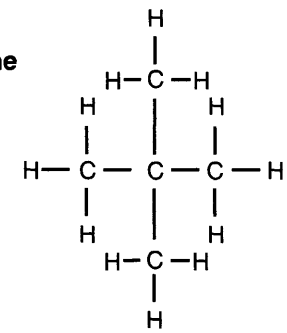
2-methylpropane



pentane



2-methylbutane



2,2-dimethylpropane

The canisters of fuel used for camping stoves or 'gas' stoves in caravans contain mainly propane and butane. Both the C_4 isomers will be present in the gas mixture.

Group presentations: *Pollution from fuels*

For this activity you will need books, leaflets and magazine cuttings about air pollution and, in particular, the pollution which results from our use of motor vehicles.

Students could work in groups, each group concentrating on a particular effect of pollution. These might include the following:

- Greenhouse effect
- Photochemical smog
- Acid rain
- Health problems caused by the toxicity of some pollutants

A number of groups could work on the same topic, or you might choose to sub-divide the topics,

e.g. health problems resulting from
- lead compounds
- carbon monoxide
- aromatic compounds

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The students should select relevant information and present it in a visually stimulating way, on a poster.

The following list of questions might help them to focus their ideas.

- What is the nature of the effect you are studying?
- Which pollutants contribute to that effect, and in what way?
- What proportion of that effect is caused by pollution from motor vehicles?

You might ask them, before the lesson, to bring in information they find for themselves in local libraries. However, it is probably safer to have plenty of resource material ready for them to use as well.

Towards the end of the lesson, when the posters are complete, you can ask each group to present the information it has collected. The posters could then be displayed so that the information is available for the class to look at in greater detail in future lessons.

Text activity 2: *How do catalytic converters help?*

Student Sheet 1.8 accompanies this activity.

Students have identified the major pollutant gases formed by petrol engines in road vehicles. They now consider what reactions could be used to convert these gases into less harmful substances.

Catalytic converters fitted into exhaust systems contain a mixture of transition metals which catalyse three main reactions. Because of this, they are known as 'three-way' converters.

The activity of the catalyst mixture is rapidly reduced by even tiny amounts of lead deposited onto it. Thus, cars fitted with catalytic converters should not be filled with ordinary, leaded, petrol. The filler pipes on most newer cars are narrower on cars fitted with converters, so that the wider filler nozzle used for unleaded fuel will not fit in.

The names of the three metals used in the catalyst mixture are given on the sheet. Students should be encouraged to locate these three elements on the Periodic Table. Bring out that they are transition metals and that these metals are frequently used as catalysts, e.g. iron in the Haber process (from the unit *Food for Thought*), nickel in hydrogenation of edible oils (from the unit *Making Use of Oil*).

Answers to questions on *Student Sheet 1.8*

Q1 Pollutants are being oxidised in reactions (i), (ii) and (iii).

Q2 Pollutants are being reduced in reaction (iii).

Q3 Catalysts change the speed of reaction but remain chemically unchanged at the end of the reaction.

Q4 Activation energy is the energy needed to enable the reaction to proceed.

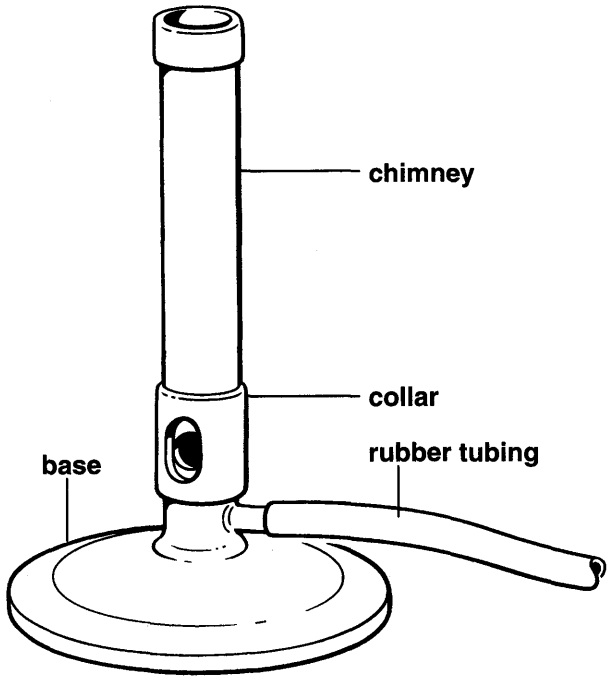
Q5 The uncatalysed reaction has the greater activation energy.

Q6 The energy released is the same in both catalysed and uncatalysed reactions.

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The Bunsen burner

A Bunsen burner can be used as a convenient example to investigate the design of gas burners.



◆ *How does it work?*

- Remove the chimney and collar and try lighting the gas from the fine jet. Describe what happens. Take care. There may be a very long flame. If the gas does not light, do not keep trying. Turn off the supply and go on to the next test.
- Try again, this time with the gas very low. Describe what happens. Then switch off the gas.
- Replace the chimney and collar, and light the Bunsen burner again. Is it easier to light with the air hole closed or open?

◆ *Investigating the controls*

- Design your own experiments to find which air hole setting gives:
  - (a) the cleanest flame
  - (b) the hottest flame.

Q1 Write a word equation and symbol equation for the complete combustion of natural gas (methane, CH_4).

Q2 Suggest two unwanted products which may form when the gas burns in a restricted supply of air.

Q3 If you had to redesign the burner to use a gas which needed a greater volume of air to burn completely, what changes might you make to the burner?

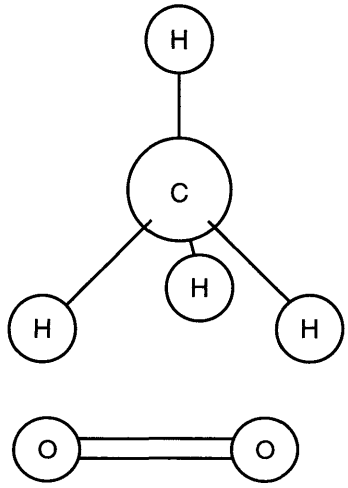


You must discuss your plans with your teacher before doing any experiments.

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Breaking and making bonds when methane burns

When natural gas burns, methane reacts with oxygen from the air. The products are carbon dioxide and water.



◆ *Making models of molecules*

In each molecule of methane, one carbon atom is bonded to four hydrogen atoms. The formula of methane is therefore CH_4.

- Make a model of a molecule of methane like this:

Carbon atoms can form **four** bonds, but hydrogen atoms can form only **one** bond.

Oxygen molecules contain two oxygen atoms.

- Make a model of a molecule of oxygen:

Each oxygen atom forms **two** bonds. Remember to use two links to join the model atoms.

◆ *Modelling chemical change*

When methane burns it reacts with oxygen. During this reaction, the bonds between the carbon and hydrogen atoms in methane must be broken. The bonds between oxygen atoms in O_2 molecules must also be broken.

- Use your models of the molecules of methane and oxygen.
- Use the models to 'break' the bonds in methane and oxygen and 'make' the bonds in carbon dioxide and water.

Q1 Carbon atoms form **four** bonds and oxygen atoms form **two** bonds. How many oxygen atoms will each carbon atom bond with?

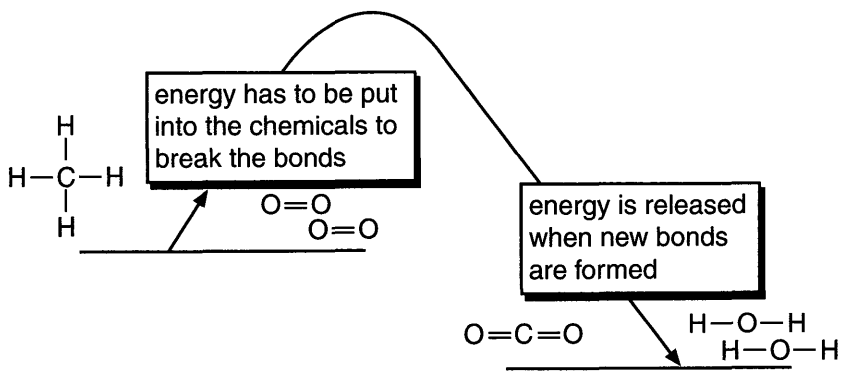
Q2 Oxygen atoms form **two** bonds and hydrogen atoms form **one** bond. How many hydrogen atoms will bond with each oxygen atom?

Q3 If all four hydrogen atoms and the carbon atom in methane bond with oxygen, how many O_2 molecules are needed to react with one molecule of methane?

◆ *Energy and chemical bonds*

Chemical bonds hold atoms together. Energy must be supplied to pull the atoms apart. Forming new chemical bonds is just the reverse. Energy is released.

When methane burns, the bonds in methane and oxygen molecules must be broken first, so that new bonds can form to make carbon dioxide and water.



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Breaking and making bonds when methane burns

◆ *Exothermic or endothermic?*

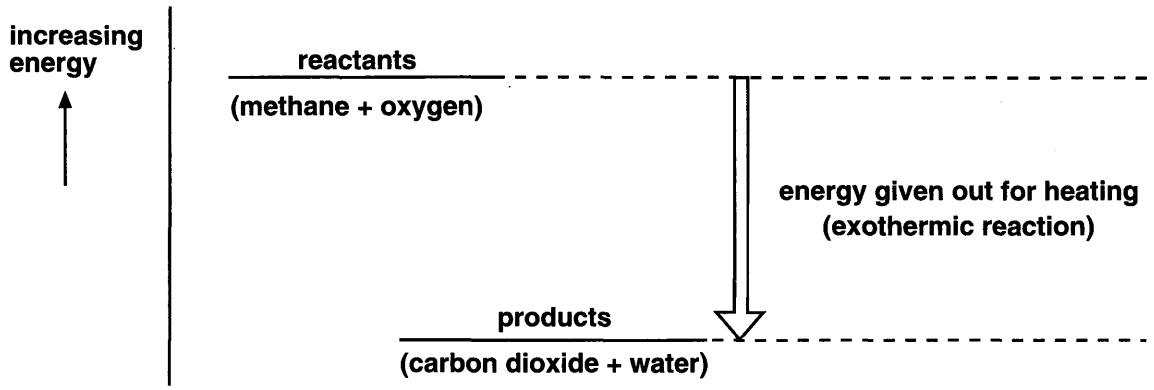
When methane reacts with oxygen, energy is given out. Most of that energy causes heating. We say that the reaction is **exothermic**.

Energy cannot be created or destroyed, only changed from one form to another. What happens when methane reacts with oxygen? Where does the energy come from?

Q4 What can you conclude about the amount of energy given out when carbon dioxide and water form compared with that needed to break the bonds in methane and oxygen?

Q5 What does this tell you about the strength of the bonds in carbon dioxide and water compared with those in methane and oxygen?

One way of representing the energy change in a reaction is by an **energy level diagram**.



In this reaction the products have less energy than the reactants. The difference in energy between the products and the reactants appears in the form of energy for heating which is lost to the surroundings.

It is possible for chemical energy to be changed into other forms of energy during a reaction, e.g. electrical energy, light or sound.

Q6 Give two other examples of reactions in which chemical energy is changed into energy for heating.

Q7 Give two examples of reactions in which chemical energy is changed into other forms of energy as well as energy for heating.

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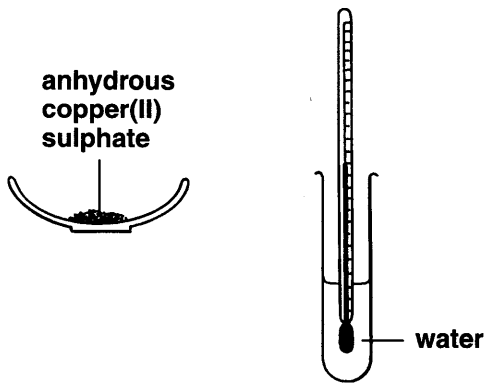
Investigating the energy transfers in reactions

When methane burns, energy is given out and the temperature of the surroundings increases. The results for this reaction (**reaction 1**) have been put in the table on *Student Sheet 1.4*.

In this experiment, you will be investigating the temperature changes and the energy transfers in some other reactions.

◆ *Reaction 2*

- Place a spatula measure of anhydrous copper(II) sulphate (**care – harmful**) on a dry watch glass and put 3 cm^3 of water in a test tube.
- Take the temperature of the water.
- Add a few drops of water to the anhydrous copper(II) sulphate.
- Cautiously feel the bottom of the watch glass.
- Put your results in the table on *Student Sheet 1.4*.

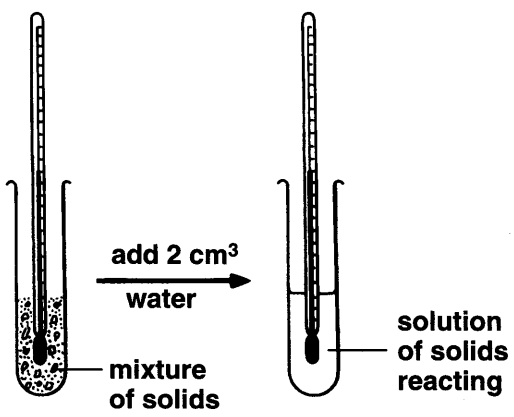


◆ *Reaction 3*

- Repeat the experiment you have just done using cement powder (**care – irritant**) or plaster of Paris instead of copper(II) sulphate.
- Put your results in the table on *Student Sheet 1.4*.

◆ *Reaction 4*

- Mix roughly equal quantities of citric acid and sodium hydrogencarbonate in a test tube (enough mixture to cover the bulb of a thermometer).
- Put the bulb of the thermometer into the mixture and note the temperature.
- Add about 2 cm^3 of the water to the mixture and note the temperature of the mixture again.
- Record the results in the table on *Student Sheet 1.4*.



Your teacher may demonstrate some further chemical reactions involving energy transfers. The results of these reactions should also be added to the table on *Student Sheet 1.4*.

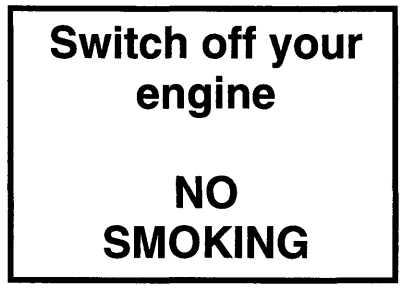
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reactants (substances at the start)	initial temperature /°C	final temperature /°C	temperature change /°C	does the temperature rise or fall?	have the substances given out energy or taken in energy?	is the change endothermic or exothermic?
methane and oxygen	20	35	+15	rise	given out energy	exothermic

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Switch off your engine

When you pull up at a petrol pump you will see this sign.



As a safety precaution you are asked to switch off your engine, as it might cause sparks, and not to smoke whilst filling up your fuel tank.

The vapour from the fuel will not burn unless there is a flame or a spark.

Investigate these two common reactions, each of which also needs a source of energy to start the reaction.

◆ *Strike a match*

- Take a match and strike it on the box. Watch carefully what happens.

Q1 Describe what happens when you strike a match.

Q2 Is energy given out or taken in when the match burns?

Q3 Do the products from the burning match contain **more energy** or **less energy** than the starting materials?

Q4 Why do the starting materials in the match head not react until the match is struck?

Q5 Why do matches have to be struck?

◆ *Light the gas*

- Turn on the gas supply to the Bunsen burner and then light the gas with a match.

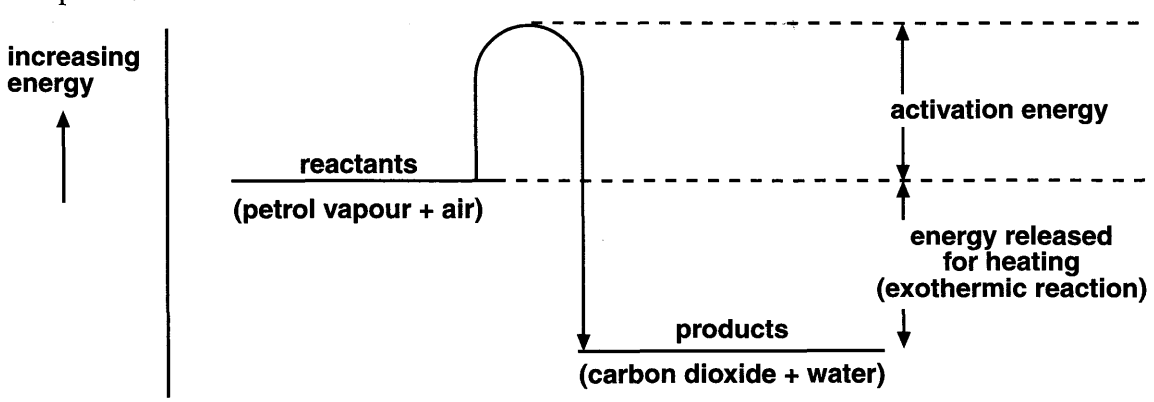
Q6 Why does the natural gas not burst into flames as soon as it mixes with the air?

Q7 Why does the natural gas burst into flames when the lighted match is brought near it?

Q8 Why does the natural gas continue to burn once it has started?

If petrol (or diesel) is spilt the vapour will not ignite spontaneously at ordinary temperatures. For a fuel to burn, bonds within the reactant molecules must be broken and new bonds made as the products form. Energy is needed to enable this bond breaking and bond making process to proceed. This energy is known as **activation energy**.

Activation energy can be shown on an energy diagram. For example, for the burning of a petrol/air mixture:



Compare this to the diagram for burning methane on *Student Sheet 1.2* (page 11).

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Flashpoints

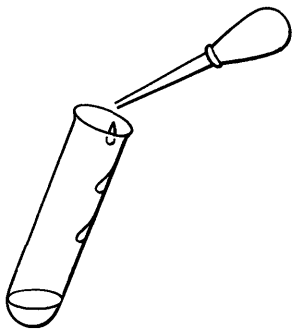
A fuel will not burn unless vapour from the fuel is present. The lowest temperature at which a liquid gives off enough vapour to burn when a flame is applied is called its **flashpoint**. This measurement allows us to compare the flammability of different fuels.

◆ *Demonstrating differences in flashpoints*

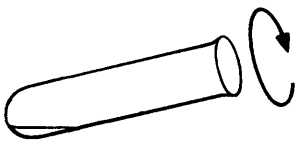


Wear eye protection.

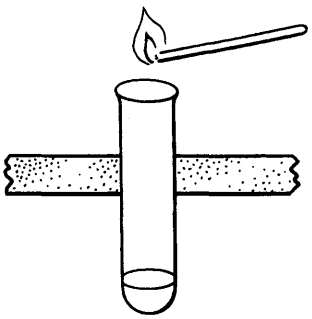
In this activity **your teacher will carry out an experiment** to compare the flammability of diesel with that of petrol.



Place half a pipette-full of petrol or diesel in a test tube.



Turn the tube round so the fuel wets the inside.



Stand the tube in a rack. Try to light the vapour at the open end of the tube

Q1 What information does your test give you about the flashpoint of diesel?

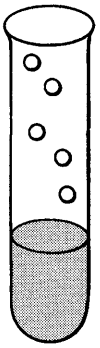
- Then, the flashpoint test will be carried out using benzene-free petrol.

Q2 What information does your test give you about the flashpoint of petrol?

◆ *Explaining the results in terms of molecules*

The molecules in a liquid are held together by attractive forces. Molecules are constantly moving and colliding with one another. As a result of a collision, a molecule may gain enough energy to overcome the attractive forces and escape as vapour.

You could draw a diagram to represent this. For example, for petrol at room temperature, enough molecules escape so that the vapour can be ignited.



- Draw similar diagrams for diesel
  - (a) at room temperature
  - (b) at its flashpoint

Q3 In which of the two fuels are the molecules held more strongly to each other?

Q4 Which of the two fuels boils at the lower temperature? Explain your answer.

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Raising the octane number

Petrol is not just one single substance. It is a mixture of different hydrocarbon molecules and additives. It is important that the fuel burns smoothly.

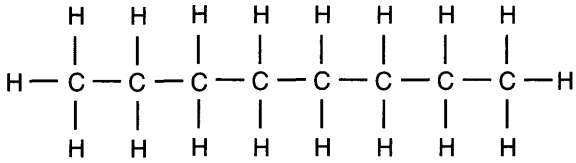
◆ *Poor quality petrol*

One of the ingredients in petrol is an alkane called **octane**.

The molecular formula of octane is C_8H_{18}.

- Make a model of an octane molecule.

This is a **straight-chain** molecule.



octane

Octane **does not** burn very smoothly.

Uneven burning of a blend of petrol because it contains ingredients which explode suddenly is called **knocking**. It reduces the efficiency of an engine.

◆ *A smoother burning compound*

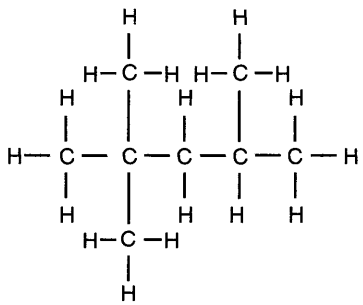
Another ingredient of petrol is 2,2,4-trimethylpentane (the old name for this was iso-octane).

The molecular formula is the **same** as for octane (C_8H_{18}).

- Make a model of this.

The atoms are arranged differently to give a **branched-chain** molecule.

2,2,4-Trimethylpentane (iso-octane) **does** burn smoothly.



2,2,4-trimethylpentane

Molecules which have the same molecular formula but which have their atoms arranged to give different molecular structures are said to be **isomers**.

◆ *A standard for fuels (the octane rating)*

Iso-octane burns smoothly and is used as a standard to which petrol blends are compared. Pure iso-octane is given a value of 100 on the octane scale. The higher the octane number the more smoothly the petrol will burn. Branched-chain alkanes have higher octane numbers than straight-chain alkanes.

◆ *How many isomers can you find?*

- Make models of molecules of each of these hydrocarbons:

  (**a**) methane, CH_4 (**b**) ethane, C_2H_6 (**c**) propane, C_3H_8

You should find that for each of these formulas there is only one possible way to arrange the atoms.

- Now see how many ways you can join atoms to make models of isomers of:

  (**d**) butane, C_4H_{10} (**e**) pentane, C_5H_{12}

- Look in chemistry books to find out what propane and butane are used for.

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How do catalytic converters help?

Very large amounts of hydrocarbon fuels are burnt in road vehicle engines. The exhaust gases can cause serious pollution problems.

◆ *The problem gases*

As well as carbon dioxide and water vapour, petrol engines also produce:

- **carbon monoxide** (which is very poisonous),
- **unburnt hydrocarbons** (poisonous, and greenhouse gases),
- **nitrogen oxides** (poisonous, contribute to acid rain).

Together, these gases can produce 'photochemical smog'.

◆ *An answer to the problem*

Car exhausts can be fitted with a **three-way catalytic converter** which contains the metals platinum, palladium and rhodium.

This helps to make the pollutants react to form less harmful gases.

(i) Carbon monoxide:

$2CO + O_2 \rightarrow 2CO_2$

(ii) Hydrocarbons:

$C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

(Any of the unburnt hydrocarbons emitted will react similarly.)

(iii) Nitrogen oxides:

$2NO + 2CO \rightarrow N_2 + 2CO_2$

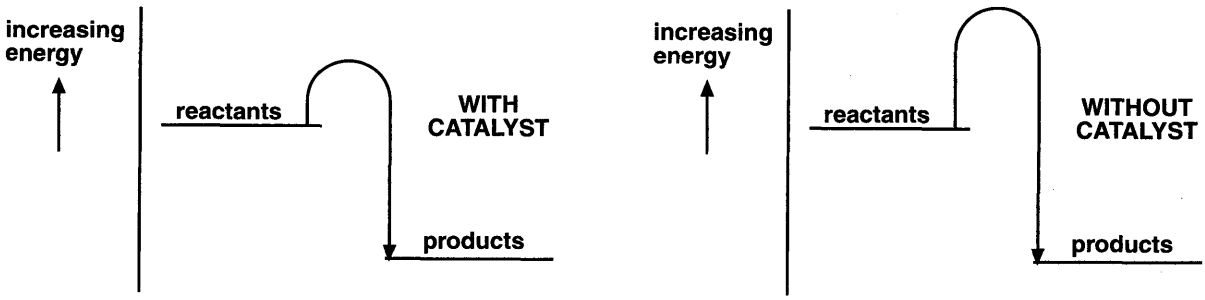
Q1 In which of the reactions are pollutants being oxidised?

Q2 In which of the reactions are pollutants being reduced?

Platinum and palladium help the oxidation reactions to occur and rhodium helps the reduction reactions to occur.

Q3 Why don't these three metals appear in the equations for the reactions?

- Look at these two energy diagrams. They describe the same reaction with and without a catalyst.



Q4 What is activation energy?

Q5 Which of the two reactions has the greater activation energy?

Q6 What can you say about the energy released in each reaction?

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Bonding

3–4 lessons

In this section students learn about how atoms can be bonded together. They may have met many of the ideas in previous work and this section acts as a way of bringing all the ideas together.

Ideas about covalent bonds are introduced using the reaction between methane and oxygen. Students' ideas about covalent bonding are extended by looking at giant covalent structures, using allotropes of carbon as examples.

Ionic bonding is introduced through discussion about corrosion, which leads to the study of ionic bonding in metal oxides. Students learn about the properties of a number of ionic compounds and find that these are related to the way that ions are arranged in giant lattices.

Ideas about the structure and bonding in metals emerge from how the physical properties of metals can be made to change.

ACTIVITY	OUTLINE	KEY POINTS
Text activity 1 (Student Sheet 2.1)	Students learn about the nature of the covalent bonds which are broken and made during the burning of methane.	Electrons are arranged in shells. A pair of electrons is shared in a covalent bond. Covalent bonding involves electrons from the outermost shell only.
Working with models 1 (Student Sheet 2.2)	Students consider models of diamond and graphite and how their different structures lead to such diverse properties.	Diamond is a 3D structure. Each carbon atom is covalently bonded to four others. Graphite is a layered structure. Each atom is bonded to three others.
Text activity 2 (Student Sheet 2.3)	Students estimate the depth at which diamonds form in the Earth's crust, and suggest suitable conditions for making synthetic diamonds.	Diamonds form at great depths in the Earth's crust. Synthetic diamonds can be made by heating and compressing graphite.
Working with models 2 (Student Sheet 2.4)	This activity about buckminsterfullerene (C_{60}) is optional but can provide interesting discussion.	C_{60} is an example of a large covalently bonded molecule.
Demonstration	Students are shown reactions of lithium, magnesium and aluminium with air and water. They discuss the implications of corrosion of these materials, which are used in the aircraft industry.	The reaction of metals with air and water is called corrosion. The principal products of corrosion are metal oxides. The properties of the oxides are different to those of the metals and to those of oxides of non-metals.
Text activity 3 (Student Sheet 2.5)	Students learn about ionic bonding in the oxides of lithium, magnesium and aluminium.	Metals react with non-metals by electron loss, forming positive ions. Some non-metal atoms gain electrons to form negative ions. The attraction between these oppositely charged ions is called an ionic bond.
Practical work 1 (Student Sheet 2.6)	Students investigate properties of ionic compounds. This leads to the idea of regularly arranged giant lattices.	Ionic compounds are brittle and have high melting points; some dissolve in water. They conduct electricity when molten or in solution.
Text activity 4 (Student Sheet 2.7)	Students apply their knowledge of ionic bonding in the context of mineralogy.	Ionic compounds are crystalline and ions are arranged in a giant lattice.
Practical work 2 (Student Sheet 2.8)	Students find the effects of applying loads of increasing magnitude to metals.	Metals, unlike most non-metals, deform elastically then plastically. They can be rolled into sheets, bent to shape or drawn into wires.
Text activity 5 (Student Sheet 2.9)	Students learn how the properties of metals can be attributed to their internal structure and bonding.	A metallic bond is the attractive force holding together metal atoms in a regular giant lattice.

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ACTIVITY	REQUIREMENTS FOR THE ACTIVITY
Text activity 1	Student Sheet 2.1 *How do atoms bond together?*
Working with models 1	Student Sheet 2.2 *Different forms of carbon* models of diamond and graphite structures
Text activity 2	Student Sheet 2.3 *Here be diamonds!* graph paper (if students **not** writing on sheet)
Working with models 2	Student Sheet 2.4 *A new form of carbon to kick around* scissors, glue
Demonstration	*Teacher demonstration:* lithium metal (**flammable and corrosive**), magnesium ribbon (**flammable**), aluminium foil 250 cm^3 beaker, Bunsen burner, metal tongs, heatproof mat
Text activity 3	Student Sheet 2.5 *Opposites attract*
Practical work 1	Student Sheet 2.6 *What are ionic compounds like?* *each group will need:* sodium chloride, potassium chloride, anhydrous sodium carbonate, aluminium oxide, magnesium oxide, anhydrous sodium sulphate Bunsen burner, test tubes, ignition tubes, test tube holders, hand lenses, 100 cm^3 beaker, connecting leads, 6V power pack, carbon rods, bulbs
Text activity 4	Student Sheet 2.7 *Some ionic minerals*
Practical work 2	Student Sheet 2.8 *A different kind of strength* *each group needs:* junior hacksaw (Eclipse saw) blades, G-clamps, stacks of 1 g and 10 g masses, Bunsen burner, metal tongs, ruler
Text activity 5	Student Sheet 2.9 *Explaining the special properties of metals* piece of galvanised iron

Student book: ***Thinking about*** *How do compounds stick together?*



Looking at *How do atoms stick together?*

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Bonding

Text activity 1: *How do atoms bond together?*

Student Sheet 2.1 accompanies this activity.

Students learn about the nature of covalent bonds and the arrangement of electrons around the nuclei of atoms.

It is advisable to work through this activity with students, as these are difficult concepts to grasp. You may wish to miss out this activity altogether with some less able groups of students.

At the beginning of the exercise the students revise that carbon forms four bonds in both methane and carbon dioxide but that, in carbon dioxide, it forms double rather than single bonds. This provides a link to Student sheet 1.2 to re-inforce basic ideas about covalent bonding.

They are reminded that atoms comprise protons, neutrons and electrons; the protons and neutrons are in the central nucleus and the electrons are found outside the nucleus. Students will be familiar with these ideas from *Seeing Inside the Body*.

Next, they are shown a table relating the number of electrons in an atom to the number of bonds formed by that atom. Carbon, for instance, has six electrons, but forms only four bonds. Students learn that electrons can be considered to occupy 'shells' and that only the electrons in the outer shell are involved in forming bonds. The first shell holds a maximum of two electrons and the second shell a maximum of eight electrons.

Helium and neon both have full outer shells. The idea that atoms bond together in such a way as to achieve the same electronic structure as these unreactive gases is introduced. A covalent bond is formed when a pair of electrons is shared between two atoms.

The students learn that outer electronic structure in atoms and in covalent molecules can be represented by means of 'dot and cross' diagrams.

Finally, the students draw some 'dot and cross' diagrams to show the covalent bonding in some other simple molecules.

Answers to questions on *Student Sheet 2.1*

Q1 Carbon forms four bonds in methane.

Q2 In methane, each hydrogen atom forms one bond.

Q3 Yes, carbon also forms four bonds in carbon dioxide.

Q4 Carbon forms two double bonds in carbon dioxide, rather than four single bonds, as in methane.

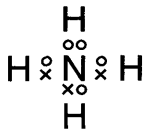
Q5 Each oxygen atom forms two bonds.

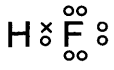
Q6 Two pairs of electrons are shared in a molecule of water.

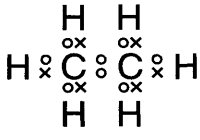
Q7 By sharing two pairs of electrons, the oxygen is now surrounded by a stable arrangement of eight electrons. Therefore oxygen forms two bonds.

Q8 Two pairs of electrons are being shared in an oxygen molecule.

Q9 Oxygen molecules contain double bonds.

Q10 (a) 

(b) 

(c) 

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Working with models 1: *Different forms of carbon*

Student Sheet 2.2 accompanies this activity.

Students have seen examples of covalent bonding between different non-metals in compounds. Atoms of non-metal elements also often bond together by forming covalent bonds.

Students could be referred to the table of electron arrangements and simple valencies from the previous activity. Discuss why noble gases such as helium and neon are the only elements which exist as single atoms (it is because the atoms already have the most stable arrangement of electrons and do not need to form bonds).

Students could also be asked to suggest how many bonds are present in a chlorine molecule (one) and a nitrogen molecule (three – with groups who respond well to abstract concepts you could link the very strong triple bond in nitrogen to its lack of reactivity as a gas, which is discussed in the units *Food for Thought* and *Evolution*).

The atoms in some elements are able to bond in two or more different ways, so that different forms of the elements can exist. Students now look at models and structural diagrams of the structures of diamond and graphite, two allotropes of carbon.

You could begin by demonstrating the marked differences in appearance and properties between graphite and diamond. You might use a real diamond, set in an engagement ring perhaps, or a diamond-headed cutter used to mark and cut laboratory glassware. Samples of synthetic diamond are available from Molehill Press, Grange Farm, Geddington, Kettering, Northants NN14 1AL.

You could show the students that diamond will scratch glass, and point out that many more diamonds are used to make cutters and drills than are cut as gemstones. Most students will know that to cut a diamond, you must use another diamond.

Students look at models and diagrams of graphite and diamond and consider how their properties are related to their structures.

For reference, the structures of graphite and diamond are included on *Student Sheet* 2.2. Point out that the greater density of diamond implies that the carbon atoms are more tightly packed than in graphite, and that this is consistent with their formation at great depths within the Earth.

Answers to questions on *Student Sheet 2.2*

Q1 Carbon dioxide.

Q2 Carbon dioxide.

Q3 Uses suggested may include:

diamond: glass cutters and drill tips (the dense, strongly bonded structure makes diamond very hard); jewellery (more difficult to relate to the structure)

graphite: in pencil leads (layers easily rubbed off to spread over the paper); in lubricants (again, layers easily slide past each other); in brushes for electric motors and in electrodes for simple cells (conductivity)

Text activity 2: *Here be diamonds!*

Student Sheet 2.3 accompanies this activity.

Student Sheet 2.3 explains why diamonds are so rare, so expensive and so hard. The sheet begins by describing the occurrence of diamonds within the Earth's crust, in kimberlite pipes.

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Students then use data to plot a graph of the conditions of temperature and pressure at which diamond turns into graphite, and vice versa. They then superimpose a graph of the conditions at different depths below the Earth's surface. From their graphs, the students estimate the minimum depth at which diamonds are likely to form. They can also suggest temperatures and pressures at which synthetic diamonds could be made in the laboratory.

Answers to questions on *Student Sheet 2.3*

Q1 24 000 atm and 760°C.

Q2 The pressure is not high enough for diamond to form at 760°C.

Q3 64 000 atm and 1200°C.

Q4 The minimum pressure needed to convert graphite to diamond at 1200°C is 42 000 atm, so diamonds **will** form at 200 km.

Q5 The minimum depth at which diamonds form is about 105 km.

Q6 You might try to make synthetic diamonds in the laboratory at the pressures and temperatures found at about 110 km, i.e. 33 000 atm pressure and 900°C.

Working with models 2: *A new form of carbon to kick around*

Student Sheet 2.4 accompanies this activity.

The discovery of buckminsterfullerene and its potential importance is described in an article by Jim Baggott in the September 1992 edition of *Chemistry Review*. Buckminsterfullerene is named after the architect R. Buckminster Fuller who designed the geodesic dome for the US Pavilion at Expo '67 in Montreal. Kroto remembered seeing the dome and eventually Richard Smalley, one of the Houston group, made the structure from 12 pentagons and 20 hexagons. If a carbon atom is placed at each junction of the seams of a soccer ball, the same structure is obtained.

A number of other fullerene molecules have now been isolated, the first being C_{70}.

The students make up a model of the C_{60} buckminsterfullerene. From the model it is easy to deduce its formula and that it has a molecular, rather than a giant, structure.

Answers to questions on *Student Sheet 2.4*

Q1 The formula of buckminsterfullerene is C_{60}.

Q2 Buckminsterfullerene, like graphite, is made of a network of carbon atoms, arranged hexagonally.

Q3 Buckminsterfullerene, unlike graphite, has a finite shape and formula.

Demonstration: *Under attack – clues to another type of bonding*

Remind students that investigations into the burning of fuels has led to an understanding of how atoms of the elements involved are held together by covalent bonds. Tell them that not all bonding can be explained by theories of covalent bonding. Investigations of reactions of metals show that there is a need for another theory of bonding.

You can start this investigation by showing them some reactions of three metals which are important in the aircraft industry. Each of these metals has a low density which makes them useful in aircraft construction.

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- aluminium – 2.70 g/cm^3
- magnesium – 1.74 g/cm^3
- lithium – 0.53 g/cm^3

Ask the students what other factors an aircraft designer should consider. Someone will probably suggest corrosion; the students may already have studied this topic in another context. If so, they will probably know that corrosion of metals is caused mainly by the action of air and water.

For all the suggested experiments, take care to follow the safety instructions.

Both teacher and students should wear eye protection. Work behind a safety screen. Use a piece of lithium no larger than a grain of rice. Tell the students not to look directly at burning magnesium.



You could demonstrate the reactions with air by holding each of the three metals in tongs in a Bunsen burner flame.

Lithium is stored in oil; before you begin, wipe off the oil using a paper towel. Do not touch lithium; instead, handle it with forceps.

Hold each metal in the flame until it begins to burn, then hold it above a heatproof mat. Lithium and magnesium burn readily leaving, in each case, a white ash. The students should know that metals react with air to form metal oxides. Aluminium does not readily burn. Explain that before they were heated all three metals had a thin coating of oxide, but that the aluminium oxide is so tough that it prevents the metal from reacting further. Aluminium, then, is a very useful metal.

You might also demonstrate the effect of water on the three metals. If you drop a lithium granule into a beaker of water, it will react immediately, dissolving and giving off hydrogen – not a desirable property for an aircraft metal! Magnesium is attacked very slowly indeed; aluminium not at all.

Clearly, aluminium is most resistant to corrosion. However, it is not sufficiently strong and hard to be much use on its own. The students may be able to suggest a solution for themselves; that is, to alloy the metals together. Alloys of magnesium and aluminium and, more recently, lithium and aluminium, are used in the aircraft industry.

At this point, you could draw attention to the fact that the oxides formed when lithium and magnesium burn in air have very different properties from those of the unreacted metals and of those of the oxides formed when fuels burn (carbon dioxide and water).

The students may themselves suggest that the oxides may have different structures and may be held together by different sorts of bonds from those which occur in metals and in the oxides of non-metals such as carbon and hydrogen. These ideas are developed later in the lesson.

Text activity 3: *Opposites attract*

Student Sheet 2.5 accompanies this activity.

The students learn how lithium and magnesium oxides are bonded together. Ionic bonding is not an easy concept to grasp, and it is intended that the teacher works through the activity with the students, adding further explanation where necessary.

Ideas about electronic structure and the stability of full outer shells of electrons have already been introduced to explain covalent bonding. They are used again to show that when metal and non-metal atoms combine, electrons are transferred so that the same kind of stability can be achieved.

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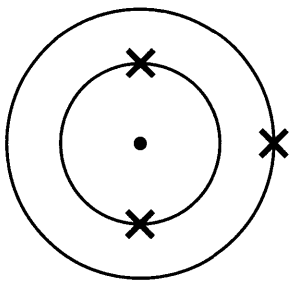
When an atom loses one or more electrons, a positive ion is formed. When an atom gains one or more electrons, a negative ion is formed. The negative ion usually changes its name: e.g. oxygen → oxide. Ionic bonds are described as the attractive forces which hold the oppositely charged ions together. The idea that ions are arranged in a giant lattice is developed in the next part of the section.

Answers to questions on *Student Sheet 2.5*

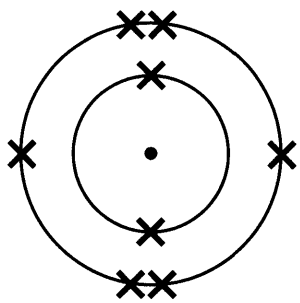
Q1 (**a**) The ratio of magnesium:oxygen atoms in magnesium oxide is 1:1.

(**b**) The ratio of lithium:oxygen atoms in lithium oxide is 2:1.

Q2 (a)

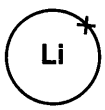


(b)

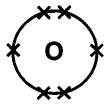


Q3 'Dot and cross' diagrams show the three types of atom as:

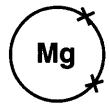
(a)



(b)



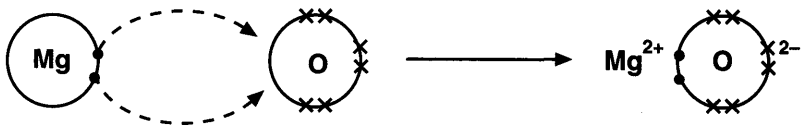
(c)



Q4 If an oxygen atom gains two outer electrons it achieves a stable outer shell of eight electrons.

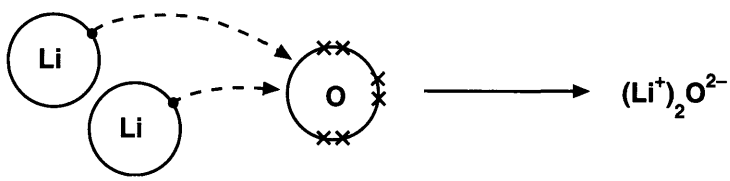
If a magnesium atom loses two outer electrons that shell becomes empty. The next shell inside is full, and therefore stable.

Q5



Q6 The magnesium ion has a +2 charge and the oxide ion, a –2 charge.

Q7



Q8 The oxygen atom needs to gain two electrons, so two atoms of lithium are needed, each to supply it with one electron.

Q9 In both magnesium oxide and lithium oxide there are equal numbers of positive and negative charges on the ions.

i.e.	Mg^{2+}	O^{2-}	overall charge
	+2	–2	0
	Li^{+}	O^{2-}	overall charge
	$2\times(+1)$	–2	0

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Practical work 1: *What are ionic compounds like?*

Student Sheet 2.6 accompanies this activity.

The students make predictions about the nature of ionic bonds by studying the properties of some ionic compounds. A selection of suitable compounds is given below.

compound	melting point / K
sodium chloride	1081
potassium chloride	1045
aluminium oxide	2313
magnesium oxide	3173
anhydrous sodium carbonate [IRRITANT]	1131
anhydrous sodium sulphate	1163

These are safe to work with, though sodium carbonate is classed as an irritant. They do not contain water of crystallisation, which would confuse melting point results. All the compounds suggested above have high melting points, though sodium chloride and potassium chloride can be melted in a Bunsen flame.

The students are asked to find out about the following properties.

- Melting point – which shows that ionic bonds are strong.
- Crystallinity – which shows that particles are arranged regularly in a three-dimensional network.
- Solubility – which shows that some ionic compounds dissolve in water and some do not.
- Conductivity – which shows that aqueous solutions of ionic compounds conduct electricity. All ionic compounds conduct electricity when molten. Ionic compounds contain charged particles.
- Brittleness – which shows that a sharp blow applied to an ionic crystal can disrupt the arrangement of alternate positive and negative ions, causing the crystal to break.

With some students you may want to lead the discussion relating the giant ionic lattice structure of ionic compounds to their high melting point and their electrical conductivity. With others you may want small groups of students to discuss their own ideas before you check their understanding. With more able students you could extend the discussion to cover why water is a good solvent for many ionic compounds.

Answers to questions on *Student Sheet 2.6*

Q1 Ionic compounds have high melting points.

Q2 Ionic bonds are strong bonds.

Q3 The uncrushed compounds contain regular-shaped crystals.

Q4 The particles in ionic compounds are arranged regularly, in a three-dimensional network.

Q5 Of the compounds suggested, sodium chloride, sodium carbonate, sodium sulphate and potassium chloride dissolve; magnesium oxide and aluminium oxide do not.

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Q6 Ionic compounds contain electrically charged particles.

Q7 In electrolysis, positive ions flow towards the cathode and negative ions flow towards the anode. When solid, the ions are held together too strongly to be able to flow.

Q8 Opposite charges attract; the ionic lattice is held together by the attractive forces between adjacent positive and negative ions.

Q9 Like charges repel; if a number of positive or negative ions were arranged together, the lattice would fall apart.

Q10 If the crystal suffered a sharp blow, part of the lattice might be displaced, causing ions of the same charge to become adjacent to one another. The crystal would break.

Text activity 4: *Some ionic minerals*

Student Sheet 2.7 accompanies this activity.

This optional extension sheet allows students to revise and consolidate ideas about ionic compounds and bonding. It could be used as a homework, or as a basis for discussion in class.

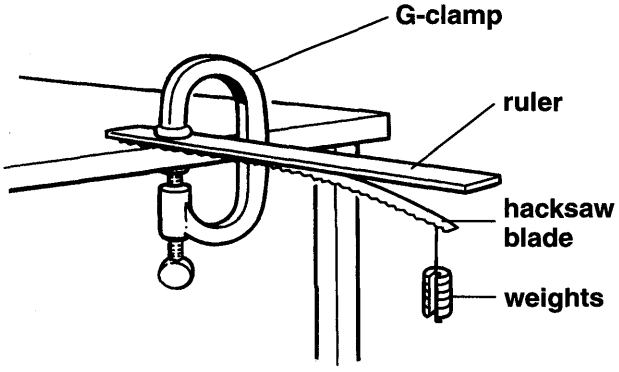
Practical work 2: *A different kind of strength*

Student Sheet 2.8 accompanies this activity.

This activity focuses on an investigation of the behaviour of metals. The intention is that the students bring enough experience of metals to the next part of the lesson to be able to contribute to a discussion about structure and bonding in metals.

Students are asked to investigate what happens when a hacksaw blade is progressively loaded with weights. The design of the experiment is left open but, in general, the hacksaw blade should be clamped firmly to the bench and weights hung from it. By removing the weights regularly, the students can check whether any deformation is elastic or plastic.

Students might try to measure the degree to which the blade must be bent before elastic deformation ceases and plastic deformation begins. They might do this by clamping a ruler on top of the hacksaw like this:



Students could use a protractor to measure the angle at which the hacksaw is bent from the horizontal.

Students should make careful observations and describe the behaviour of the blade as accurately as they can, using their newly learnt vocabulary.

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The weights are exerting a bending force on the blade. As loading begins, the blade deforms elastically. The more weights that are added, the greater the deformation. At one point, the blade will fail to return to its original position; it has now begun to deform plastically. It is likely that, eventually, the blade will fail by bending very sharply. **Warn the students to be careful that the weights do not drop on their feet.** You could tell them to put a rubbish bin or bucket containing waste paper on the floor beneath the weights, so that they are forced to stand well out of harm's way. If the blade is heated and then cooled quickly its properties change. Very few weights will be needed before the blade fractures. Sudden cooling causes metals to become brittle. This is explained further in the next part of the lesson.

If a piece of metal is continually bent and unbent, it will become increasingly stiff. The students try doing this using a paper clip or a length of thick copper wire.

The students should appreciate that metals can be made to behave in a wide variety of ways and that they can be very strong, yet flexible.

Answers to questions on *Student Sheet 2.8*

Q1 Hacksaw blades are made of steel.

Q2 They are hard, which gives them a long life, but they are strong and flexible, so they do not snap under stress.

Q3 The hacksaw blade fractures under a very small load.

Q4 Bending or shaping the metal causes it to harden, hence the name.

Text activity 5: *Explaining the special properties of metals*

Student Sheet 2.9 accompanies this activity.

In this part of the lesson students learn how the physical properties of metals can be explained by their internal structure. By now, they should know that metals, on the whole, have high melting points. They will have been reminded of this when they heated the steel hacksaw blade. They should also know that metals are good conductors of heat and electricity, and that they are strong, yet flexible. Most metals have high densities.

You can demonstrate that the metals are crystalline by showing the students a piece of galvanised iron. They may recall that this is iron that has been dipped into molten zinc. Crystal boundaries can be seen quite clearly in the zinc coating.

Alternatively, you could cast some zinc by melting it in a crucible over a Bunsen flame and pouring it into a depression made in a tray of damp sand. Do this behind a safety screen, or preferably in a fume cupboard. Once the nugget has cooled, you can crack it open with a hammer, again behind a safety screen; inside, there will be a crystalline mass of solid zinc.

From the data on the information sheet, students should be able to suggest that, in general, metals have high melting points, high densities and good electrical conductivity.

Common properties suggest that the bonding is similar in all metals; it must be different from the covalent bonding in diamond because diamond does not conduct electricity.

How can metals be so flexible when they are held together by such strong bonds? And how can metals conduct electricity? Explain to students that metal atoms have small numbers of electrons in their outer shells. They tend to lose these electrons in an attempt to become more like the noble gases (see Text Activity 3, *Student Sheet 2.5*).

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Within the metal lattice there are charged ions surrounded by loosely attached electrons which can be made to move around when an electric potential is applied. The attractive forces between the positively charged ions and the negatively charged electrons hold the lattice together.

A model that is often used is to regard the lattice of metal ions as permeated by the loosely attached electrons. When a piece of metal is bent, lines of ions in one part of the crystal are pushed across lines of ions in another part of the crystal. While this is happening, the ions are still held together because the electrons still permeate the structure. To move lines of electrons in a diamond crystal, covalent bonds would have to be broken and could not be reformed until the atoms had been put back in a new arrangement in which covalent bonds could reform. The structure would snap before this could happen.

Student Sheet 2.9 provides a simplified summary of these ideas.

Answers to questions on *Student Sheet 2.9*

Q1 Answer to involve movement of electrons.

Q2 Any suitable answers.

Q3 Too expensive.

Q4 Light/low density.

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How do atoms bond together?

You should be very familiar with the following reaction:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

You have learned that during the reaction, bonds are broken and new bonds are made.

What are these bonds?

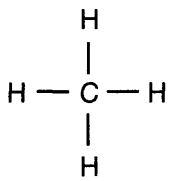
What causes a carbon atom and four hydrogen atoms to join together in a molecule of methane?

This is a detective puzzle which you can solve by studying the evidence accumulated by scientists who have studied the behaviour of atoms and molecules.

◆ *How many bonds?*

Let's begin with methane.

Analysis shows its formula to be CH_4 and its structure to be:



Q1 How many bonds does the carbon atom form in methane?

Q2 How many bonds does each hydrogen atom form in methane?

Look at the structures of carbon dioxide and water.

O=C=O H—O—H

Q3 Does carbon form the same number of bonds in carbon dioxide as it does in methane?

Q4 How do the bonds formed by carbon in carbon dioxide differ from those in methane?

Q5 How many bonds does each oxygen atom form?

What holds the atoms together in simple molecules like methane, carbon dioxide and water?

To answer this question you need first to check your understanding of the structures of atoms.

◆ *What particles are atoms made of?*

Atoms are thought to contain three different kinds of particles:

- protons, which are positively charged;
- neutrons, which have no charge;
- electrons, which are negatively charged.

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How do atoms bond together?

◆ *Atomic structure*

Experiments on thin sheets of metal led to the idea that protons and neutrons form the central nucleus of an atom, and that the electrons move continuously around the nucleus.

The attractive force between the positively charged nucleus and the negatively charged electrons holds the atom together.

In a chemical reaction the nucleus of the atom is unchanged. The breaking and making of bonds involves only the electrons.

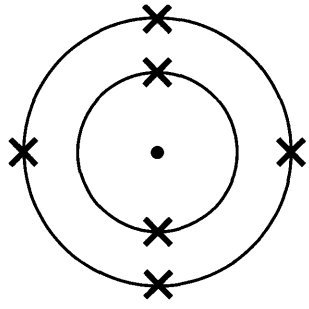
Are all the electrons involved, or just some?

Look at the table below, which gives information about the first ten elements of the Periodic Table.

atom	number of electrons	number of bonds formed in simple molecules	notes
HYDROGEN	1	1	
HELIUM	2	0	helium is very unreactive
LITHIUM	3	1	
BERYLLIUM	4	2	
BORON	5	3	
CARBON	6	4	
NITROGEN	7	3	
OXYGEN	8	2	
FLUORINE	9	1	
NEON	10	0	neon is very unreactive

Carbon has six electrons but forms only four bonds.

To explain this, and the behaviour of the other atoms, we consider the electrons to be arranged in 'shells' around the nucleus. The shells extend all around the atom, but on a flat page we draw them as circles.



e.g. for carbon

Two of the electrons are in the inner shell.

Only the four electrons in the outer shell are used in forming the four bonds.

- Try drawing a diagram like this for oxygen.

You might expect oxygen to form six bonds, but it usually forms only two.

To explain this, we need to look at atoms like helium and neon which are very unreactive. Their arrangements of electrons are very difficult to change.

Scientists think that other atoms react together to form compounds so as to achieve similar arrangements of electrons.

- Draw diagrams to show the arrangement of electrons in an atom of helium and an atom of neon.

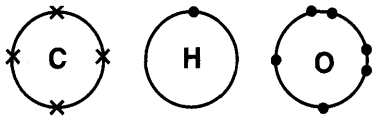
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How do atoms bond together?

◆ *Bonding in simple molecules*

When drawing diagrams to show how atoms form bonds, you can use the symbol for the element to represent the nucleus, and draw small dots or crosses to represent electrons round it.

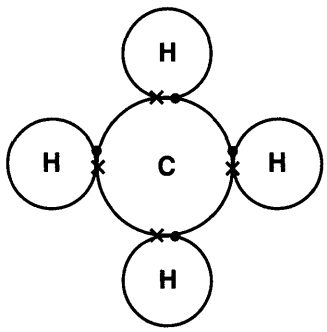
Since only the electrons in the outer shell form bonds, we often show atoms in the following way:



Each bond between hydrogen and carbon has a **pair** of electrons which are shared by both atoms.

Each hydrogen atom has its own electron (which it shares with the carbon atom) plus the one which it shares from the carbon atom. In other words each hydrogen atom has a share in two electrons – the number of electrons found in the very stable, unreactive helium atom.

We can represent methane like this.



The carbon atom is surrounded by its own four electrons plus four more electrons, one shared from each hydrogen atom. In other words, the carbon has eight electrons round its outer layer – the 'neon arrangement'.

The carbon and hydrogen atoms are bonded together in a methane molecule because they are sharing four pairs of electrons. This is known as **covalent bonding**. Each shared pair of electrons is a **single covalent bond**.

- Draw a 'dot and cross' diagram to show how electrons are shared in a molecule of water, H_2O.

Q6 How many pairs of electrons are being shared in one molecule of water?

Q7 Explain why oxygen forms two bonds rather than six bonds.

A 'dot and cross' diagram of carbon dioxide looks like this.



The carbon atom shares two pairs of electrons with each oxygen atom.

Carbon dioxide contains **two double covalent bonds**.

- Draw a 'dot and cross' diagram to represent the bonding in a molecule of oxygen, O_2.

Q8 How many pairs of electrons are being shared in a molecule of oxygen?

Q9 Are the two oxygen atoms held together by single or double bonds?

Q10 Draw 'dot and cross' diagrams of the covalent bonding in molecules of:

(a) ammonia, NH_3;

(b) hydrogen fluoride, HF;

(c) ethane, C_2H_6.

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Different forms of carbon

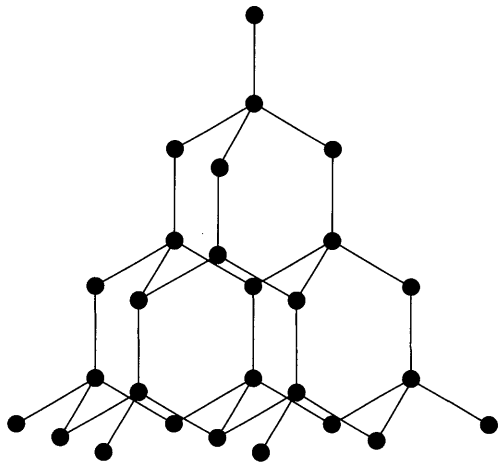
In most non-metal elements the atoms are held together by covalent bonds. Sometimes it is possible for a particular type of atom to bond together in two (or more) different ways.

This gives different structures for the element. The different forms are called **allotropes**.

Carbon can form different allotropes.

◆ *Diamond*

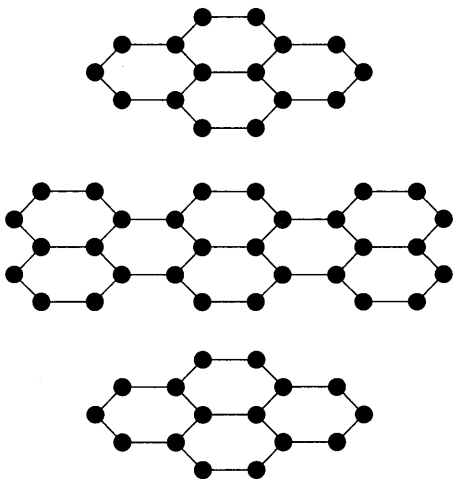
In diamond each carbon atom is covalently bonded to four other carbon atoms. A single diamond contains millions and millions of carbon atoms held together in this way. This is a very strong structure so that diamond is very hard and has a high melting point (3550°C).



The density of diamond is 3.53 g/cm^3.

Q1 What would you expect to be formed if you heated a diamond very strongly in air?

◆ *Graphite*



In graphite each carbon atom is covalently bonded to three other carbon atoms. This results in layers of atoms in which the carbon atoms are arranged in a hexagonal pattern.

Three of the outer electrons of each carbon atom are used to form the three bonds. The fourth electron is not fixed in place. This is why graphite can conduct electricity, whereas diamond cannot.

The bonding forces holding the layers together are weak so they are able to slide over each other. This is why graphite is flaky and slippery.

The density of graphite, 2.25 g/cm^3, is lower than diamond because the atoms are less tightly packed.

Q2 What would you expect to form if you heated graphite strongly in air?

Q3 Try to find out some of the uses of diamond and of graphite. For each one, explain how the structure gives it suitable properties for the use.

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Here be diamonds!

Until about a hundred years ago, diamonds had only been found in river gravels.

Diamonds, eroded from the rocks, had been transported by streams and rivers, and finally deposited in river valleys.

◆ *Where are diamonds found?*

In 1870, some diamonds from a South African river gravel were traced back to their source rock near what is now the city of Kimberley.

The source of the diamonds was a deep volcanic vent made of an unusual looking rock which became known as **kimberlite**.

As the miners dug down the pipe grew narrower, but still it yielded diamonds – the 'Big Hole' at Kimberley is the deepest man-made pit in the world!

Since then, many more of these vents, called **kimberlite pipes**, have been discovered in Siberia, North America and in other parts of Africa.

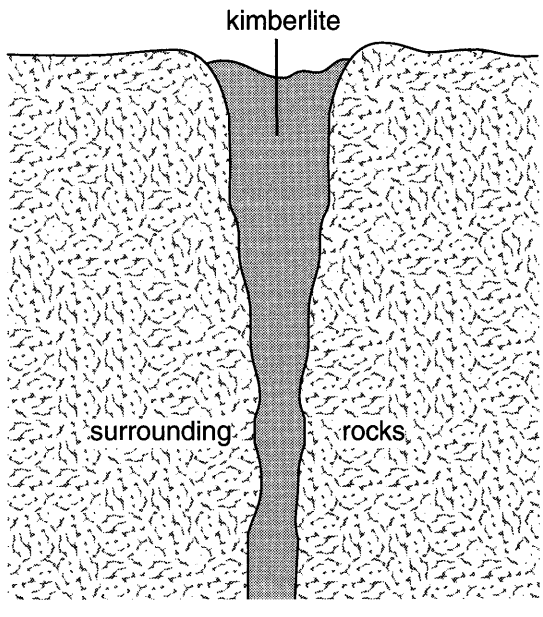
This is a cross-section of a kimberlite pipe.

Not every kimberlite pipe has diamonds in it.

Geologists believe that diamonds form at great depths. They are carried upwards in molten lava which solidifies to form kimberlite rock.

Nobody has ever excavated down to the base of a kimberlite pipe.

Scientists use information about temperatures and pressures at different depths in the Earth's crust to estimate the depth at which diamonds form.



◆ *Where do diamonds form?*

The figures given in this table are estimated using data collected from deep mines.

depth/km	pressure /atm × 1000	temperature/°C	notes
50	16	600	The deeper you go down a mine, the greater the pressure exerted on the rocks.
100	32	880	
150	48	1060	
200	64	1200	The deeper you go down a mine, the higher the temperature.
250	80	1330	
300	96	1460	

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Here be diamonds!

◆ *A graph to predict diamond formation*

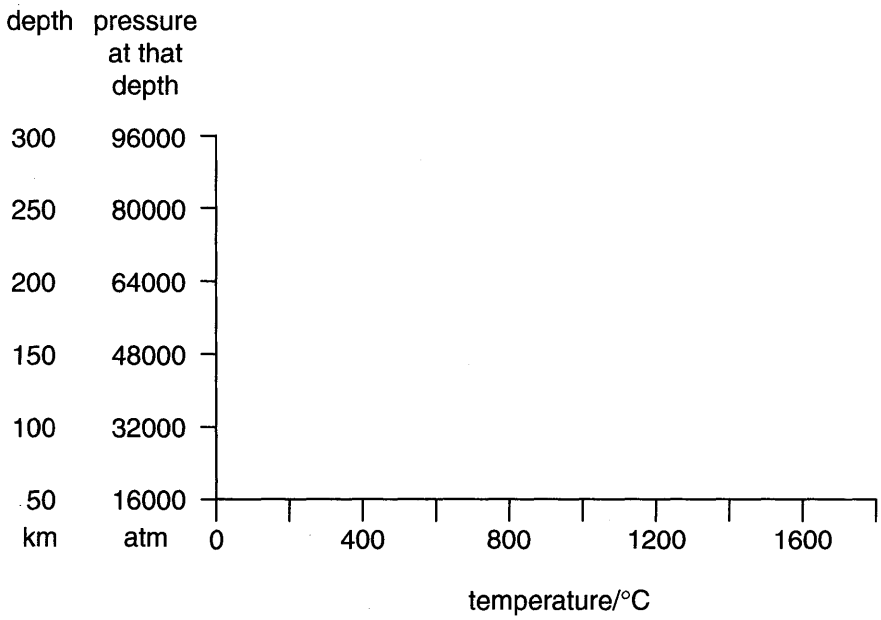
Diamond and graphite are both forms of the element carbon.

It needs very high temperatures and pressures to form diamond.

The figures in the table show the temperatures and pressures at which the transition from graphite to diamond occurs.

temperature /°C	pressure /atm × 1000
600	26
800	32
1000	36
1200	42
1400	48
1600	53
1800	59

- Draw a graph, using axes like these.



- On your graph, plot the information about the temperature and pressure needed to form diamond.
- Mark the area above the line 'diamond formed'.
- Mark the area below the line 'graphite formed'.
- Now add a second line to your graph, using information from page 34, about the temperature at each depth below the surface.

Q1 What is the pressure and temperature at a depth of 80 km?

Q2 Why will diamonds not form at this depth?

Q3 What is the pressure and temperature at a depth of 200 km?

Q4 Explain how the conditions make it possible for diamonds to form at a depth of 200 km.

Q5 What is the minimum depth at which diamonds will form?

There have been many attempts to make synthetic diamonds; though some successes were reported in the nineteenth century, all were frauds. A method for making synthetic diamonds was first discovered in the 1950s.

Q6 Suggest conditions which might give diamonds in a laboratory experiment.

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A new form of carbon to kick around



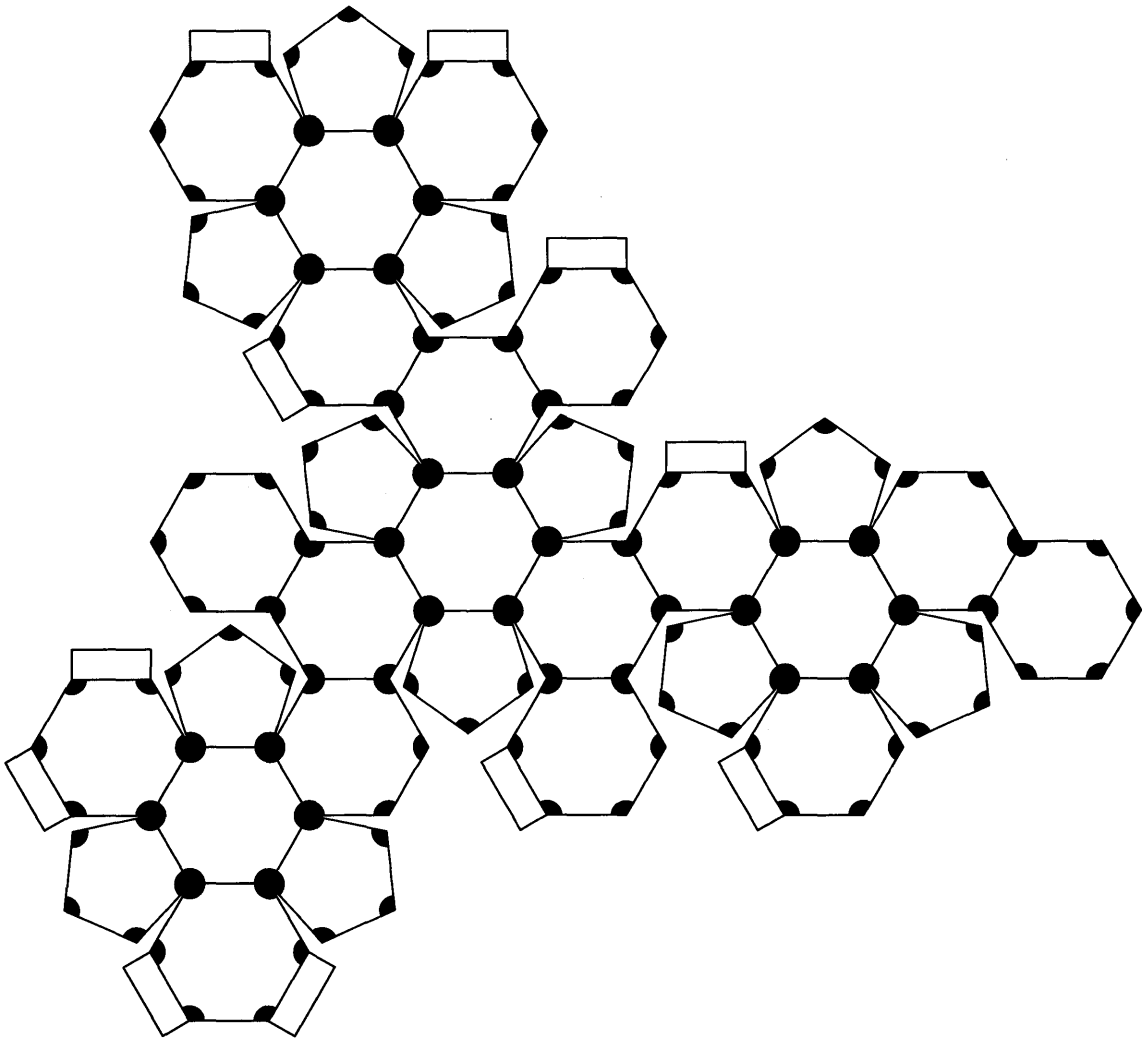
A new form of carbon, called **buckminsterfullerene,** was discovered in 1985.

Harold Kroto and his group of scientists from the University of Sussex, collaborated with a group at the University of Houston, Texas.

They had been trying to demonstrate that long-chain carbon compounds could be formed from graphite by subjecting it to the conditions similar to those found on certain stars.

To their surprise they found that they had made a carbon molecule which had a shape like a soccer ball. They made the first model of the structure from paper – like the one below!

These scientists were awarded the Nobel prize in 1996 for their work on buckminsterfullerene.



- Make up this model of buckminsterfullerene by cutting round the outline, folding all the edges and sticking down the tabs.

Q1 What is the formula of buckminsterfullerene?

Q2 In what ways does buckminsterfullerene resemble graphite?

Q3 In what ways does buckminsterfullerene differ from graphite?

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Opposites attract

◆ *Burning metals*

Magnesium (a metal) burns readily in oxygen (a non-metal). Opposites, it seems, attract.

The compound formed is magnesium oxide. It is quite easy to analyse its composition and find that its formula is MgO.

Lithium also forms an oxide: its formula is Li_2O.

Q1 What is the ratio of metal atoms to oxygen atoms in:

(a) magnesium oxide;

(b) lithium oxide?

The two oxides have similar properties. They contain the same type of chemical bond. Why do their formulas differ? To answer this question we need to look at the atomic structures of the three elements.

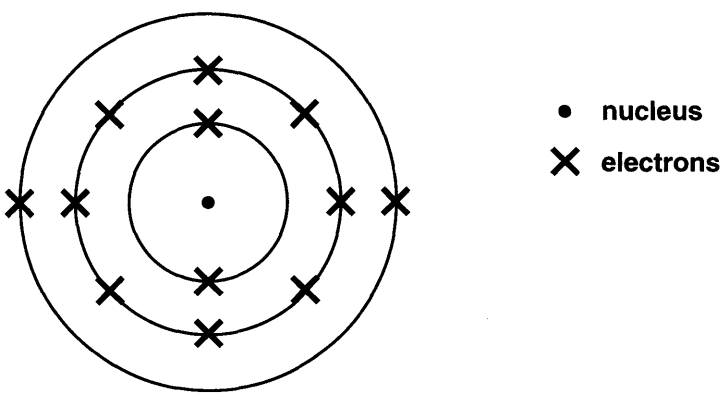
Lithium	**Magnesium**	**Oxygen**
3 protons	12 protons	8 protons
4 neutrons	12 neutrons	8 neutrons
3 electrons	12 electrons	8 electrons

◆ *How are particles arranged in atoms?*

You probably remember that the **protons** and **neutrons** lie in the central **nucleus** of the atom, and that the **electrons** move around the nucleus.

You may also remember that the electrons are arranged in **shells** and that after the first shell, which holds a maximum of two electrons, the others can hold up to eight electrons.

A magnesium atom can therefore be represented as:



Q2 Draw similar diagrams for atoms of:

(a) lithium;

(b) oxygen.

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Opposites attract

◆ *How are atoms of the elements held together?*

Q3 Draw 'dot and cross' diagrams showing the outer electron arrangements for atoms of:

(a) lithium;

(b) oxygen;

(c) magnesium.

These atoms can achieve full outer shells by losing or gaining electrons. In magnesium oxide, each magnesium atom has lost two electrons and each oxygen atom has gained two electrons.

Q4 Explain how both types of atom gain full outer shells by doing this.

Atoms contain equal numbers of protons and electrons, so they have no overall electrical charge. When electrons are lost or gained electrically charged particles, or **ions**, are formed.

Q5 Draw a 'dot and cross' diagram to show how electrons are transferred from magnesium to oxygen when magnesium oxide is formed.

Q6 What are the charges on each type of ion?

Opposite electrical charges attract one another. Two oppositely charged ions are held close to one another by an attractive electrical force. This is called an **ionic bond**.

Q7 Draw a 'dot and cross' diagram to show how electrons are transferred from lithium to oxygen in lithium oxide. The ions formed have the formulas Li^+ and O^{2-}.

Q8 Explain why the formula of lithium oxide is Li_2O.

Q9 Explain why lithium oxide and magnesium oxide have no overall electrical charge.

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What are ionic compounds like?

The properties of ionic compounds can tell us a lot about the nature of the ionic bonds that hold them together.

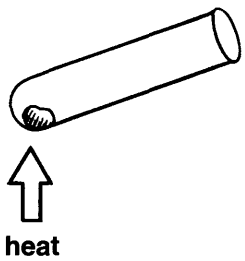
Your teacher will give you a selection of ionic compounds to work with. At first sight they may not appear to have much in common, but, if you study them carefully, you will find that many of them share certain properties. See what patterns you can find.

◆ *Melting point*

- Heat a small amount of substance in an ignition tube, at first gently, then strongly.

Q1 What can you conclude about the melting points of ionic compounds?

Q2 What does this suggest about the strength of ionic bonds?





Wear eye protection.

◆ *Crystallinity*

- Many of the compounds you will use have been crushed into powders. However, some are uncrushed.
- Look at samples of these under a hand lens or microscope.

Q3 What do you see?

Q4 What does this tell you about the way that the particles are arranged in ionic compounds?

◆ *Solubility in water*

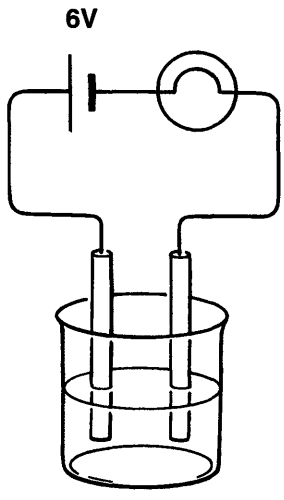
- Try dissolving a very small amount of substance in a test tube half-filled with water.

Q5 Which compounds dissolve and which do not?

◆ *Electrical conductivity*

- Try passing an electric current through solutions of some ionic compounds, using a 6 V battery and graphite electrodes. Include a lamp in your circuit to indicate whether a current is flowing.

Q6 What do your results tell you about the nature of the particles present in ionic compounds?



All ionic compounds conduct electricity when they are molten.

Q7 Why do you think ionic compounds conduct electricity in the liquid state, but not in the solid state?

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What are ionic compounds like?

◆ *Brittleness*

- Try crushing the uncrushed samples using a pestle and mortar. Some ionic compounds are hard, but all of them are brittle.

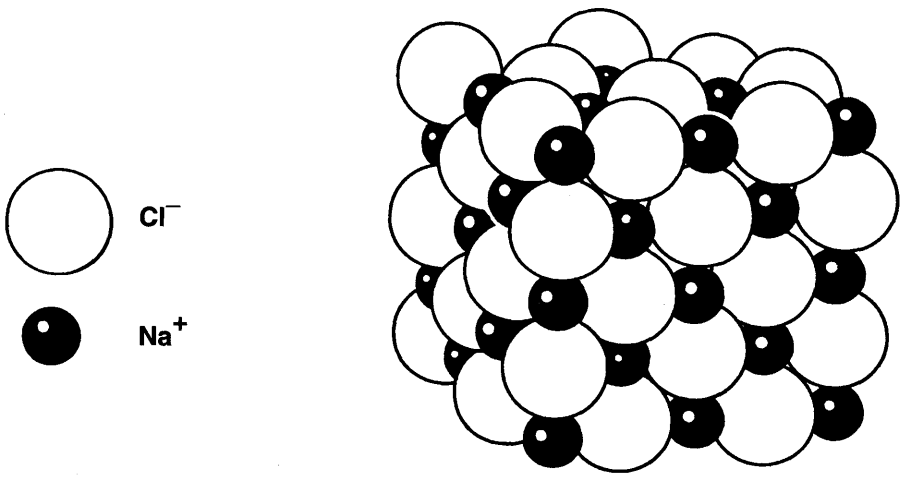
◆ *Explaining the properties of ionic compounds*

When sodium reacts with chlorine, oppositely charged ions, Na^+ and Cl^-, are formed. The salt which is formed is not at all like the element it was from from.

Oppositely charged ions attract each other so that when millions and millions of them come together a crystal of salt is formed.

In the crystal the ions are arranged in a particular pattern and the structure is called a **giant ionic lattice**.

Part of the lattice for sodium chloride is shown below. The ions in all ionic compounds are arranged in giant ionic lattices.



Q8 Explain why the positive and negative ions are arranged alternately.

Q9 What would be the effect on the lattice if a number of positive or negative ions were all arranged together?

Q10 What might happen to the arrangement of ions in the lattice if the crystal was hit with a hammer? How does this explain the brittleness of ionic compounds?

- Discuss how the giant ionic lattice structure of sodium chloride explains its high melting point and that it conducts electricity when molten.

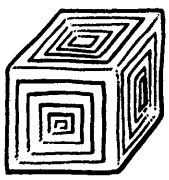
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Some ionic minerals

Many minerals are ionic compounds. Look at these extracts from an old mineralogy textbook, then try the activities. 'H' is hardness on Mohs scale, numbered from talc (1), the softest, to diamond (10), the hardest.

SP. GR. is specific gravity, the density compared to the density of water.

ROCK SALT, HALITE, Common Salt
COMP. Sodium chloride, NaCl; calcium chloride and magnesium chloride are usually present, sometimes also calcium sulphate and magnesium sulphate; the presence of magnesium compounds causes the mineral to become wet and lumpy. CRYST. SYST. Cubic. COMMON FORM. Commonly cubes as shown below.



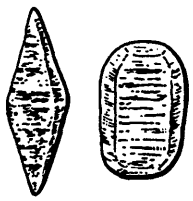
COLOUR. Colourless or white when pure; often yellow or red, sometimes bluish. LUSTRE. Vitreous; transparent to translucent. TASTE. Saline, soluble in water. FRACT. Brittle. H. 2–2.5. SP. GR. 2.2.

USES. Rock salt is used for culinary and preserving purposes, and in a number of chemical manufacturing processes.

SYLVINE, Sylvite.
COMP. Potassium chloride, KCl. CRYST. SYST. Cubic. COMMON FORM. Cubic, massive or granular. COLOUR. Colourless or white. LUSTRE. Vitreous. TASTE. Saline, more bitter than that of rock salt. Soluble in water. H. 2. SP.GR. 2.

FLUORSPAR, FLUORITE, Blue John
COMP. Calcium fluoride, CaF_2.
CRYST. SYST. Cubic. COMMON FORM. Crystals of cubes common; also massive and granular. COLOUR. Colourless, white, green, purple, blue or yellow. LUSTRE. Vitreous, transparent to translucent. FRACT. Uneven, brittle. H. 4. SP. GR. 3–3.25.

CORUNDUM
COMP. Aluminium oxide, Al_2O_3. CRYST. SYST. Hexagonal-trigonal. COMMON FORM. Barrel-shaped or pyramidal crystals, as shown below.

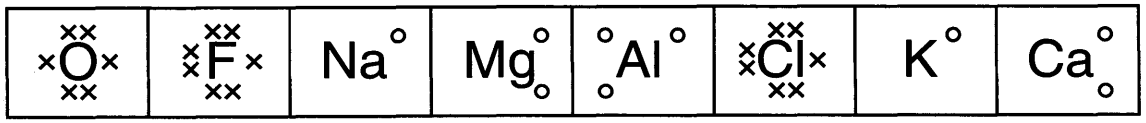


COLOUR. Of common varieties, grey, greenish and reddish. More rarely, *ruby* (red) and *sapphire* (blue). LUSTRE. Vitreous. FRACT. Uneven. H. 9 (next to diamond). SP. GR. 3.9–4.1.

CRYOLITE
COMP. Fluoride of aluminium and sodium. CRYST. SYST. Monoclinic. COMMON FORM. Crystals not common, usually found massive. COLOUR. Colourless or white, reddish or brownish. LUSTRE. Vitreous; subtransparent to subtranslucent. FRACT. Uneven, brittle. H. 2.5. SP. GR. 2.97.

USES. Cryolite is used in the manufacture of aluminium, being added to the aluminium electrolyte to lower its melting temperature.

- List the pieces of evidence that suggest that these minerals are ionic compounds.
- The outer electron arrangements of atoms of the elements mentioned in the extracts are shown below.



- Draw 'dot and cross' diagrams to show how ions form in the minerals sylvine, fluorite and corundum.
- Halite (mainly sodium chloride) can also contain magnesium chloride and calcium chloride. Work out the formulas for these two impurities.

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A different kind of strength

There are three ways in which materials might respond if put under stress (that is, if a force is applied).

ELASTIC DEFORMATION – the material deforms, but springs back once the force has been removed.

PLASTIC DEFORMATION – the material deforms, and remains changed in shape after the force has been removed.

FRACTURE – the material breaks.

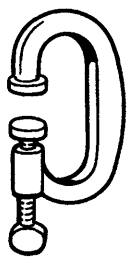
Is there any link between the way particles are arranged in a substance and the way it reacts to stress?

◆ *Metals under stress*

Use the following apparatus to observe carefully what happens when a hacksaw blade is gradually loaded with an increasing number of weights.



a hacksaw blade



a G-clamp



some weights

Your method should allow you to determine, at any stage, whether the deformation is elastic or plastic.

- Write down your observations.

Q1 What metal is the hacksaw blade made of?

Q2 How does the response of that metal to stress make it ideal for its purpose?

In fact, the blade has been treated by careful heating and cooling so as to improve its properties.

You can undo that process by holding the blade in tongs and heating it until it is red hot, and then cooling it quickly in a bowl of cold water.

- Do this, and then try loading the blade with weights once again.

Q3 How does the behaviour of the hacksaw blade differ from before?

Metals are truly remarkable materials. They can be strong and hard, but very flexible. Unlike strong, non-metallic materials such as diamond, they can be bent into shape, drawn into wires and beaten into thin sheets. Within limits, the same metal can be brittle or bendy, depending on the treatment it receives.

- Try bending and unbending a piece of thick copper wire, or a straightened out paper clip, over and over again. What do you notice?

Q4 Suggest why this process is given the name 'work-hardening'.

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Explaining the special properties of metals

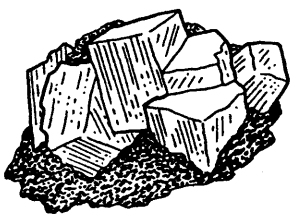
Metals are strong, yet flexible. They can be beaten into sheets, bent into shapes and drawn into wires. Molten metals can be poured into moulds and cast into shapes.

Some of the properties of some metals are shown in the table below.

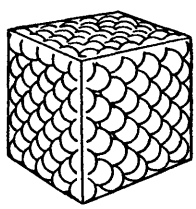
metal	melting point/°C	density/g/cm^3	thermal and electrical conductivity	order of conductivity (1 is highest)
iron	1539	7.86	good	8
aluminium	659	2.70	good	4
zinc	420	7.13	good	6
lithium	181	0.53	good	7
lead	328	11.35	good	9
gold	1063	19.32	good	3
magnesium	650	1.74	good	5
copper	1083	8.94	good	2
silver	961	10.50	good	1

◆ *Bonding in metals*

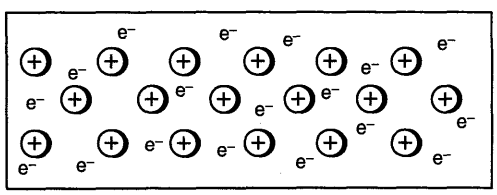
The special properties of metals can be explained in terms of their structure and bonding.



Metals can form regularly shaped crystals.



In a metal crystal, the atoms are arranged in a giant lattice.



Metal atoms lose electrons easily to form positive ions. In a giant metallic lattice, the positive metal ions are surrounded by a 'sea' of loosely attached electrons. The attractions between positively charged ions and negatively charged electrons are known as metallic bonds.

Q1 Explain how their structure enables metals to conduct electricity.

Q2 Find uses for four metals in the list and explain how their typical metallic properties enable them to do these jobs.

Q3 Why do we not use silver for the bases of saucepans even though it is a better conductor than copper?

Q4 Why is aluminium used for overhead power cables?

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A chemical family: the halogens

Photography provides a context for study of the elements that make up the halogen family.

2–3 lessons

Students study the silver halides and the bleaching action of the halogen elements. The students learn that the halogens show marked similarities in some of their properties.

Students learn how photographic film and paper are made and how they work. Given data on chlorine, bromine and iodine, students are asked to make predictions about the nature of fluorine and astatine.

ACTIVITY	OUTLINE	KEY POINTS
Discussion (Student Sheet 3.1)	A discussion about photography leads to the idea that certain substances are sensitive to light.	Silver chloride, bromide and iodide are sensitive to light. They are used in photography. These silver halides are all insoluble in water.
Practical work 1 (Student Sheet 3.2)	Students prepare and purify samples of silver chloride, bromide and iodide. They compare the appearance of these salts, and their reaction when exposed to light.	Insoluble salts can be prepared by ionic precipitation. Silver halide salts darken on exposure to light because they react to form metallic silver.
Practical work 2 (Student Sheet 3.3)	Students use photographic paper to make shadow prints which they can develop and fix to produce a permanent image.	Light-sensitive photographic paper is made by coating paper with a layer of one or more silver halide salts.
Homework suggestion 1	Students find out how photographic film and paper are made, and how an image is developed and fixed.	
Demonstration	Some properties of the halogens are demonstrated.	The halogen elements occupy Group VII of the Periodic Table. They are all volatile non-metals, and have similar properties.
Practical work 3 (Student Sheet 3.4)	Students investigate the bleaching action of the halogens. They convert a black and white photograph to a 'sepia' print, using iodine solution.	Aqueous chlorine, bromine and iodine react in a similar way with silver in black and white photographs. Aqueous chlorine and aqueous bromine are both bleaching agents. Iodine is a less powerful oxidiser.
Data activity (Student Sheet 3.5)	From information about chlorine, bromine and iodine, students predict the nature of fluorine and astatine.	The elements of Group VII have similar properties because the elements have the same number of outer electrons.
Homework suggestion 2 (Student Sheet 3.6)	Students read and answer questions about the properties and uses of the halogens, as described in prose and poetry.	

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ACTIVITY	REQUIREMENTS FOR THE ACTIVITY
Discussion	Student Sheet 3.1 *Images in silver* camera which is not loaded with film reel of developed film some black and white photographs
Practical work 1	Student Sheet 3.2 *Silver halides* *0.1 mol/dm^3 solutions of the following reagents:* silver nitrate, potassium chloride, bromide and iodide apparatus for filtering, wash bottles, stop clocks, lamps
Practical work 2	Student Sheet 3.3 *Making shadow prints* access to laboratory with black-out or efficient blinds selection of objects which cast interesting shadows medium-grade photographic paper **check the labels for hazards** commercial photographic developer **check the labels for hazards** commercial photographic fixer **check the labels for hazards** newspapers or kitchen towel, shallow plastic trays forceps or tongs, stop clocks, lamps
Demonstration	*The halogens* well-stoppered tubes containing chlorine gas (**toxic**), liquid bromine (**corrosive and very toxic**) and iodine crystals (**harmful**) a few iodine crystals in a boiling tube sealed with cling film beaker of water, at about 35°C
Practical work 3	Student Sheet 3.4 *Now you see it . . .* access to fume cupboard black and white photographs, colour photographs scissors aqueous solutions of equivalent concentrations of chlorine, bromine and iodine (for details, see Teachers' Notes) access to clocks or watches with seconds hands fine paintbrushes, plastic tubs or trays, forceps dilute iodine solution alkaline sodium sulphide solution (**corrosive**) (for details, see Teachers' Notes) access to a 1M solution of sodium thiosulphate
Data activity	Student Sheet 3.5 *Close-up on the halogens*
Homework suggestion 2	Student Sheet 3.6 *A host of uses*

Student book



Looking at *Photographs* provides an additional homework.

Thinking about *4. The periodic table – how does it help?* revises key ideas.

Things to do Questions 2 and 10 provide additional homework.

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A chemical family: the halogens

Students carry out experiments to illustrate the main chemical processes in black and white photography. The properties of chlorine, bromine and iodine are demonstrated and students explore uses of the halogens as bleaches.

A data interpretation exercise is used to reinforce understanding of family relationships amongst the halogens.

Discussion: *Images in silver*

Student Sheet 3.1 accompanies this activity.

You could begin the lesson by asking the students how they think black and white photographs are made. Some students may have photography as a hobby and may know about developing and printing; most will have, at some time, taken photographs. You might prompt the discussion by showing the students a camera which is not loaded with film, a reel of film which has been developed and some black and white photographs. The following points should arise from the discussion.

- A camera is loaded with photographic film which contains substances that are sensitive to light.
- When the shutter is opened, the camera lens focuses light from the subject onto the photographic film (the camera was studied in the unit *Communicating Information*).
- Where light falls, the film darkens and a **negative** image is formed.
- Once finished, the film is developed.
- Black and white photographs are made by shining light through the negative onto photographic paper which is also light sensitive. A **positive** print is made.

The discussion could then turn to the nature of the light-sensitive materials used in photography. As early as 1727, it was known that some silver compounds darken when they are exposed to light, but photography became possible in the early nineteenth century when the salts, silver chloride, bromide and iodide, were found to react very quickly to light. Explain that these three salts are known collectively as the silver halides and that they were first prepared by reacting the elements together.

e.g. silver + chlorine → silver chloride

The students should know that chlorine is toxic and unpleasant to handle. Explain that bromine and iodine are similarly noxious, and that they can try more convenient ways of preparing the silver halides starting from compounds rather than elements.

Answers to questions on *Student Sheet 3.1*

Q1 Photographic emulsion: the layer of light-sensitive material on the film.

Negative image: the light intensity of the image produced on the film is the reverse of that in the view which is photographed. Light-coloured objects reflect most light onto the film, causing most chemical change and producing the darkest patches on the film. Dark objects reflect less light, so causing less change and leaving the film clear.

Developers: chemicals which will continue the conversion of silver halides to silver. These reactions are catalysed by silver metal. The greater the light intensity at any point in the film, the greater the amount of silver which will have been produced by the light-sensitive reaction, so the faster the developer reaction will go.

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Places on the film where most light has fallen darken fastest (tiny specks of silver block the passage of light). Students may be reminded of the action of catalysts, and that silver is a transition metal.

Fixers: chemicals which dissolve away any unreacted silver halides, so that no further reaction can occur.

Positive photographic print: black and white photographic paper is coated with silver halides. Dark patches on the negative block the passage of light, so the corresponding points on the print will be lightest in colour. The light/dark register, which was reversed on the negative, is restored to match that of the original scene.

Q2 (a) Some of the crystals should have dark dots at the centre.

(b) The crystals which had dark dots are now the centre of dark patches.

(c) The dark patches remain, but the crystals between them, which were not affected by light, have gone.

Practical work 1: *Silver halides*

Student Sheet 3.2 accompanies this activity.

Warn the students that silver salts are corrosive and will cause brown stains if spilt on the skin. These stains can be removed by washing the affected area with a solution of sodium thiosulphate of concentration 1 mol/dm^3.

The students make samples of insoluble silver chloride, silver bromide and silver iodide by ionic precipitation. For example, when making silver chloride, aqueous silver ions (in the form of silver nitrate solution) are mixed with aqueous chloride ions (for instance, in the form of potassium chloride solution).

$$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$$

In fact, any soluble silver salt and any chloride may be used; all that is necessary is that aqueous silver ions mix with aqueous chloride ions.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

This type of reaction is called ionic precipitation.

The silver halides can be separated from the mixture by filtering; they form the residue which may be purified by washing. If this preparation is done quickly it is possible to note the appearance of the precipitate before it reacts to light, and to note the time it takes for the silver halide to darken. The appearance and behaviour of each of the three halides can then be compared. If it is a cloudy day, or the laboratory does not receive much sunlight, it may be advisable to shine light from a lamp directly onto the precipitate. An overhead projector lamp works very well indeed.

Students should be encouraged to think of ways in which a fair comparison of the three silver halides can be made. Silver chloride darkens most quickly, then silver bromide; silver iodide does not darken at all. The students learn that the dark colour is due to the formation of metallic silver. If they find it difficult to believe that this dull, grey substance is silver, you could show them that the shiny crystals which form on a spiral of copper wire, dipped into a test tube containing aqueous silver nitrate, disintegrate into a powdery, grey mess when the wire is taken out of the solution.

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Answers to questions on *Student Sheet 3.2*

Q1 silver nitrate + potassium chloride → silver chloride + potassium nitrate
silver nitrate + potassium bromide → silver bromide + potassium nitrate
silver nitrate + potassium iodide → silver iodide + potassium nitrate

Q2 $AgNO_3 + KCl \rightarrow AgCl + KNO_3$
$AgNO_3 + KBr \rightarrow AgBr + KNO_3$
$AgNO_3 + KI \rightarrow AgI + KNO_3$

Q3 (a) potassium K^+ chloride Cl^-
(b) potassium K^+ bromide Br^-
(c) potassium K^+ iodide I^-
(d) silver Ag^+ nitrate NO_3^-

Q4 $Ag^+ + Cl^- \rightarrow AgCl$
$Ag^+ + Br^- \rightarrow AgBr$
$Ag^+ + I^- \rightarrow AgI$

Practical work 2: *Making shadow prints*

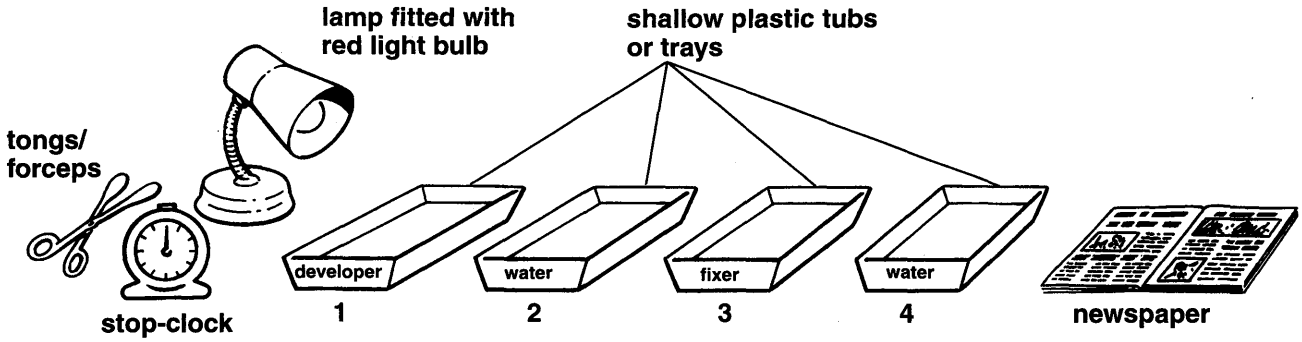
Student Sheet 3.3 accompanies this activity.

The students make shadow prints using commercial photographic materials. This sounds expensive and difficult to do without darkroom facilities but, if planned carefully, good results can be achieved, first time, in a laboratory with blackout or reasonably efficient blinds. It is great fun, and a good example of how the light-sensitive nature of the silver halides can be put to use!

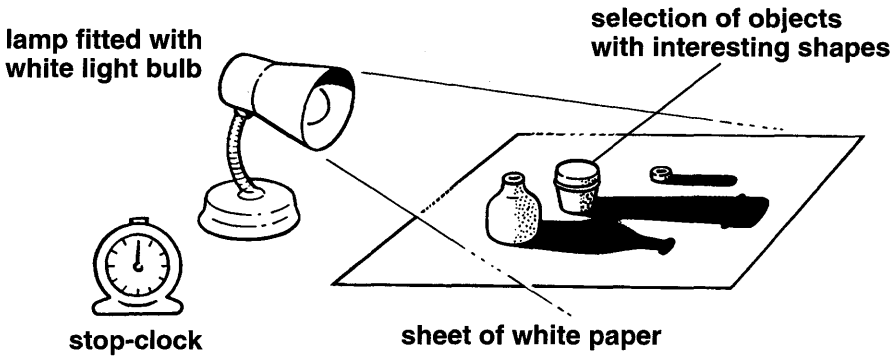
Medium-grade resin paper works well and is available from shops selling photographic equipment. To save expense, it can be cut into small pieces before the lesson. This should be done in a darkroom, using a red safelight. Store the pieces of paper, wrapped in aluminium foil, in a black polythene bag until they are needed.

Developer and fixer are also available commercially. Follow the manufacturers' instructions carefully; these will show you how to dilute the chemicals and how long the prints take to develop and fix. Be sure to read and take heed of the hazard warnings issued by the manufacturer: students must wear eye protection and gloves.

It is probably best to set up several processing stations around the laboratory. At each one there should be the following:



Set up stations at which the shadow prints can be made. At each one, there should be the following:



Make the laboratory as dark as possible.

The students might first try casting shadows onto the ordinary white paper. They can then be given a piece of photographic paper on which to make the shadow print.

The shorter the exposure time, the greater will be the contrast between the light and dark areas of the print. To get the best results the students could experiment using a range of exposure times, but a reasonable print should be achieved by illuminating the photographic paper for about ten seconds.

Once the objects have been removed, the photographic paper must be protected from white light until the image has been developed and fixed. A good place to store the undeveloped prints would be between the pages of a book.

When all the shadow prints have been made, the white lights should be turned off and the red safelights turned on. The students can then, in turn, develop and fix their shadow prints by placing them first into developer (for 1–2 minutes, depending on the exposure time), then into fixer and finally into water. A water bath placed between the developer and fixer baths will help prevent contamination and will allow one pair of tongs or forceps to be used for the whole operation. The shadow prints can then be left on newspaper to dry.

Homework suggestion 1: *How are photographic film and paper made?*

Students can find out how photographic film and paper are made and how an image is developed and fixed.

Demonstration: *The halogens*

The rest of this section is about the chemistry of the halogen elements. A few people are exceptionally sensitive to halogens; they usually know this, because the sensitivity will have become apparent at swimming pools. It is advisable to check beforehand if there are any students who experience such problems, as it may be necessary for them to stay out of the laboratory during the work on halogens.

In the previous activity the students will have investigated some of the properties of silver halide salts, and will have noticed similarities between the compounds. You might begin by asking the students to suggest from which elements chloride, bromide and iodide salts are formed. Point out the positions of chlorine, bromine and iodine on the Periodic Table and suggest that their appearance in the same group might give them, like the groups of metallic elements they have studied in previous units, some degree of similarity. In short, that these elements may show some family likenesses.

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When asked what they know about the appearance of chlorine, bromine and iodine, students sometimes give confused answers. They often say that chlorine is a liquid, because many swimming pool disinfectants are in liquid form. These are commonly referred to simply as 'chlorine'. Students may be under the impression that iodine is a brown liquid because laboratory solutions of iodine in potassium iodide are brown. They will have encountered bromine water in *Food for Thought* and *Making Use of Oil*, but have probably not seen pure bromine before. Have ready to show the students three well-stoppered tubes containing small samples of chlorine gas (**toxic**), liquid bromine (**corrosive and very toxic**) and solid iodine (**harmful**).

The students should not handle containers of chlorine, bromine and iodine.

Keep the containers sealed.



It is easy to see that bromine is volatile because above the brown liquid will be reddish-brown vapour. Iodine vapour above the iodine crystals can be seen if the tube is viewed end-on.

You might show the purple colour of iodine vapour more clearly by immersing a test tube containing a few crystals of iodine, and sealed with cling film, into a beaker of warm water at about 35°C. Remove the tube from the water as soon as purple vapour becomes visible. The iodine vapour will quickly recondense as small crystals on the side of the tube.

Do not allow iodine vapour to escape from the tube.



Ask the students whether these elements are metallic or non-metallic. Chlorine and bromine are clearly non-metals. Iodine crystals have a metallic lustre, but their volatility suggests that iodine, too, is a non-metal.

On first sight a pale green gas, a brown liquid and some silver-grey crystals might seem to bear scant resemblance to one another.

Explain to the students that in the rest of this section they will learn about some of the similarities and differences in the properties of this group of elements, which are collectively called the Group VII elements.

Tell the students that this group was formerly, and sometimes still is, called the halogens, and that this word means 'salt-giver'. Since they have already encountered halide salts, they should be able to suggest why this is an appropriate name.

It is important to stress the difference between the terms halogen and halide. For instance, chlorine is a toxic, greenish-coloured gas, whereas sodium chloride, the ionic compound formed when sodium and chlorine react together, is an unreactive, white solid. One common misconception is that it is unsafe to add **fluoride** to drinking water supplies because **fluorine** is poisonous!

In the next activity students use aqueous solutions of chlorine, bromine and iodine. You might at this stage show them the appearance of these three solutions: aqueous chlorine is almost colourless, aqueous bromine is a deep orange colour, iodine dissolved in potassium iodide is brown.

Stress that even aqueous solutions of the halogens are toxic, and that they should be kept and used in a fume cupboard.

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Practical work 3: *Now you see it . . .*

Student Sheet 3.4 accompanies this activity.

The students investigate some of the properties of aqueous solutions of chlorine, bromine and iodine.

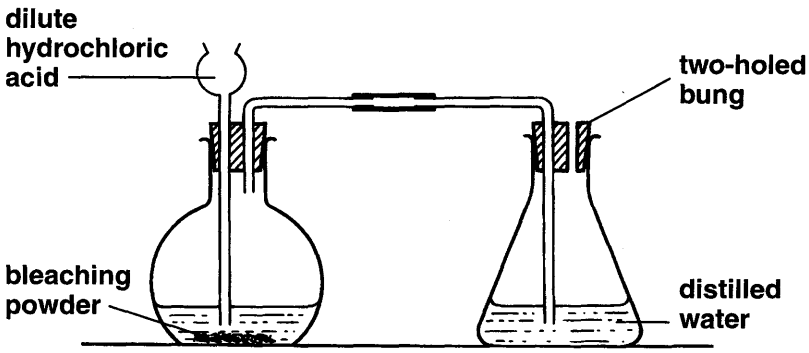
These solutions should be made up and, at all times kept, in a fume cupboard.



For your information

To allow a fair comparison, it is important that the three solutions have equivalent concentrations.

Chlorine is not very soluble in water, so saturated aqueous chlorine solution is used. It should be made up, in a fume cupboard, as shown in the diagram below.



Allow chlorine gas to bubble into water for about 15 minutes. The aqueous chlorine solution thus formed should be pale green in colour and should readily bleach litmus paper.

The most effective way of dealing with spills of chlorine, bromine or iodine solutions is by using a solution of sodium thiosulphate of concentration 1 mol/dm^3.

A beaker containing this solution should be placed in the fume cupboard for use in emergencies.

Work in a fume cupboard. Wear gloves and a face shield whilst handling pure bromine. Do not inhale bromine vapour.



To prepare bromine water of an equivalent concentration, make up about 50 cm^3 of saturated aqueous bromine by adding liquid bromine drop by drop to 50 cm^3 of water, with mixing, until some undissolved bromine remains at the bottom of the beaker.

Decant off 37 cm^3 of the saturated aqueous bromine into a measuring cylinder and make up to 100 cm^3 with distilled water. The remains of the saturated solution may be diluted down and the resulting solution used generally in the laboratory.

Iodine is not very soluble in water but dissolves readily in potassium iodide, giving a brown solution. For these experiments, make up a solution of potassium iodide of concentration 1 mol/dm^3 by dissolving 16.6 g of solid potassium iodide in 100 cm^3 of distilled water. Then add to this solution 2.16 g of iodine crystals and stir until they, too, have dissolved.

Liquid bromine is difficult to handle. The beaker should be as close as possible to the bromine stock bottle in a fume cupboard. All the aqueous halogens should be stored in stoppered bottles.



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The students are going to try, using the solutions of chlorine, bromine and iodine, to remove silver from black and white photographs. You could use strips cut from the shadow prints they made in the previous lesson, though commercially developed photographs work better. You might ask the students to bring in some unwanted black and white and colour photographs. In another part of the activity, students investigate the bleaching action of the halogen solutions on the dyes in colour photographs.

This activity must be carefully managed, because the experiments will have to be carried out in a fume cupboard, which students will have to queue up to use. However, strips of photographs can be cut elsewhere in the laboratory and the time taken to do the experiment is short.

The students are asked to dip strips of black and white photographs into each of the three solutions for 30 seconds and then to compare the results. Aqueous chlorine, bromine and iodine all react with the silver in black and white photographs, rendering it colourless, but they do so at different rates; iodine reacts most quickly, chlorine least quickly. Once the strips of photograph have been washed, they can be taken to another part of the laboratory for scrutiny.

The students are asked which of the solutions they would choose if they wanted to remove a small part of the photograph. Aqueous chlorine solution is not very effective and, though the iodine solution removes the silver the most quickly, it leaves behind a deep brown stain. Aqueous bromine works best, and students may wish to try removing part of a photograph using a fine paintbrush. This should be done in a fume cupboard, and care should be taken that the aqueous bromine solution does not touch the skin.

The students then investigate the bleaching action of the three solutions, first by dipping in strips of colour photographs for 30 seconds, washing them and comparing the results. Aqueous chlorine and bromine solutions will both remove the blue and green dyes with about equal efficiency, but will not remove red or yellow dyes. Iodine solution does not bleach the dyes at all.

Finally, students should test the bleaching action using some other dyes. They could dip a piece of blue litmus paper into each solution, and add a drop of blue ink to each solution. Again, aqueous chlorine and bromine solutions bleach the dyes, but the iodine solution does not.

One interesting use of iodine is in the production of 'sepia' prints from black and white ones, and you might allow students to do this, if time allows.

It is vital to work in a fume cupboard. Dilute some of the iodine solution ten times and place it in a plastic tray or tub to about 1 cm depth. Pour into another plastic tray, to the same depth, some distilled water and into a third tray, some alkaline sodium sulphide solution (**corrosive**), made by mixing equal volumes of 0.1 mol/dm^3 sodium sulphide and 0.5 mol/dm^3 sodium hydroxide solution. The addition of alkali does not impair the quality of the result and prevents hydrolysis of sodium sulphide to form hydrogen sulphide. Dip the black and white photograph into the iodine solution until it becomes faded and yellow. This will take from 30 seconds to 1 minute. Wash the print in water and then place it in the tray containing the sodium sulphide solution. After about 1 minute the yellow colour will have disappeared and a stronger image appeared, but in 'sepia' tones rather than black and white. The print can then be washed and left to dry.

Answers to questions on *Student Sheet 3.4*

Q1 Bromine solution works best; chlorine solution works slowly and, whilst the reaction with iodine is rapid, a brown stain is left on the photograph.

Q2 Solutions of chlorine and bromine bleach the dyes from coloured photographs; a solution of iodine does not.

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Q3 Blues and greens are affected the most, reds and yellows the least.

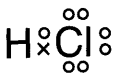
Q4 Solutions of chlorine and bromine will bleach blue litmus and blue ink; a solution of iodine will not.

Data activity: *Close-up on the halogens*

Student Sheet 3.5 accompanies this activity.

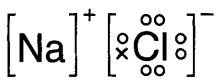
The students are given information about the properties and atomic structure of chlorine, bromine and iodine, and asked to predict the properties of two other halogens, fluorine and astatine. The purpose of the activity is to highlight some of the similarities between the elements of Group VII and to suggest some explanations for them. By the end of the exercise it is hoped that the students will appreciate that the halogens and their compounds bear a striking family resemblance to one another, and that there is a gradation of properties as the group is descended. All the halogens possess seven outermost electrons and need one more electron to achieve a complete shell. They form covalent bonds with non-metallic elements.

e.g.



With metals, the halogens form ionic bonds.

e.g.



You can use these similarities in electronic structure and in the types of bond formed to give a simple explanation of why the elements of the halogen family have similar properties. Point out that atoms of different halogens have, for instance, different radii, so their properties are not identical.

Answers to predictions on *Student Sheet 3.5*

Appearance

Q1 Fluorine is a gas at room temperature. Astatine is a solid at room temperature. Silver fluorine is a white solid, silver astatine is a yellow solid.

Chemical reactions

Bromine reacts with metals more vigorously than iodine but less vigorously than chlorine. Silver bromide reacts at a moderate rate with light.

Answers to questions on *Student Sheet 3.5*

Q1 Chlorine is the most reactive, iodine the least reactive.

Q2 Fluorine is more reactive than chlorine.

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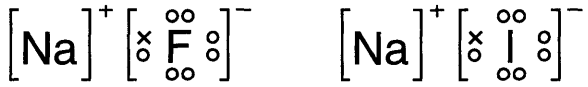
Q3 The formulas are F_2, At_2, NaF, NaAt

The 'dot and cross' diagrams for chlorine and iodine are as follows:



Q4 The atoms of chlorine and iodine are achieving a full shell by sharing a pair of outer electrons.

The 'dot and cross' diagrams for sodium fluoride and sodium iodide are as follows:



Q5 The atoms of fluorine and iodine are both gaining an electron from an atom of sodium. The negatively charged fluoride and iodide ions thus formed have complete outer shells of electrons.

Q6 The family resemblance of the halogens can be attributed to their each having the same number of outer electrons.

Homework suggestion 2: *A host of uses*

Student Sheet 3.6 is needed for this text-related activity.

The students find out about the uses of the halogens and their compounds by reading extracts of prose and poetry.

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Images in silver

A sample of silver chloride, silver bromide or silver iodide (silver halides) will get darker in colour if it is exposed to light.

Light energy causes a reaction which produces silver metal.

◆ *What is film made of?*

Photographic film has a layer of silver halide spread over a transparent material such as cellulose acetate. Early photographs used glass plates in place of film.

To make it spread and stick more evenly the silver halide is mixed with gelatine to form **photographic emulsion**.

Gelatine swells and softens when it is wetted, so the chemicals used in processing can reach the silver halide more easily.

◆ *Developing and fixing*

Our ancestors look very serious in photographs. They had to sit very still because they had to sit for a long time to let enough light into the camera to develop a **negative image**.

Snapshots became possible because of the invention of chemicals called **developers**.

When light falls onto a silver halide crystal a tiny part of that crystal may react to form metallic silver. You can't see any dark colour, but a record of the scene, a latent image, has been formed. When the film is processed the developer turns much more of the crystal to silver and a full image develops.

There is no need to expose the film for more than a fraction of a second in most light conditions. Action photographs become possible.

If you expose a film to too much light, all the silver halide crystals will begin to react to form silver. When you develop the film it will be black all over. So film must be kept away from light; many modern films are packed in plastic cassettes so that you can load and unload your camera in the light.

Why then, can you look at processed negatives in ordinary light? After developing, the image is fixed by washing it in a solution which dissolves away the unreacted silver halide, so that no more of the crystals can react to form silver. These chemicals are called **fixers**.

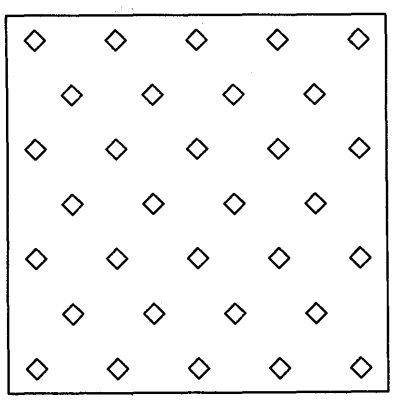
◆ *Making a print*

If you shine a light through a negative and focus it onto light-sensitive photographic paper, a **positive photographic print** will be formed. This is then developed and fixed just like the negative film.

Q1 Use the information in the text to explain the photographic terms in **bold print**.

Q2 The diagram represents a piece of photographic film coated with silver halide crystals. Use similar diagrams to represent what happens when the film is

(a) exposed to light for a short time;
(b) developed;
(c) fixed.



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Silver halides

The silver salts used in photography are silver chloride, silver bromide and silver iodide. Together, these are known as **silver halides**. The silver halides are insoluble in water.

◆ *Making a sample of the silver halides*

The silver part is provided by using **silver nitrate solution**; this contains **silver ions**.

The halide part is provided by using:

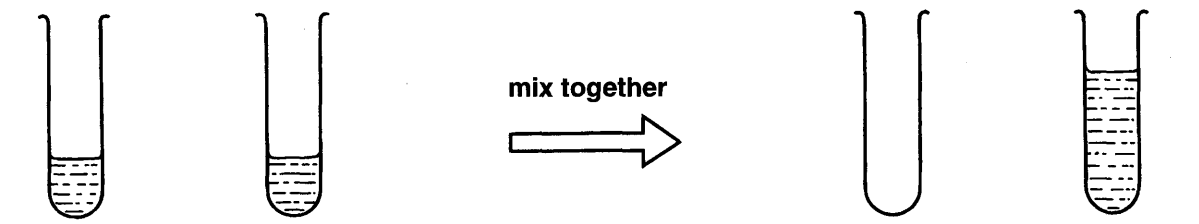
potassium chloride solution to make silver chloride;
potassium bromide solution to make silver bromide;
potassium iodide solution to make silver iodide.

When made this way the silver halides appear as solids.



Wear eye protection.

◆ *Making silver chloride*



- Keep the test tube which contains the mixed solutions.

◆ *Making silver bromide and silver iodide*

- Repeat the experiment, first using potassium bromide and then using potassium iodide instead of potassium chloride.

◆ *Recording the observations*

Copy and complete the following table.

chemicals mixed	silver halide made	appearance of the solid
silver nitrate + potassium chloride	silver chloride	
silver nitrate + potassium bromide		
silver nitrate + potassium iodide		

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Silver halides

◆ *What chemical reactions form silver halides?*

Q1 Write a word equation for the formation of each of the silver halides you have made.

Q2 Write a symbol equation for the formation of each of the silver halides you have made.

Q3 Write the name and symbol for the ions present in each of the following:

(a) potassium chloride;

(b) potassium bromide;

(c) potassium iodide;

(d) silver nitrate.

Q4 Write an ionic equation for the formation of each of the silver halides you have made.

◆ *Are the silver halides affected differently by light?*

Plan an investigation which will enable you to answer this question.

In each of your three test tubes you have a solid silver halide mixed with a solution of another chemical.

When planning your investigation you need to consider the following:

- on a sunny day the natural light from the Sun can be used;
- a fair comparison needs to be carried out;
- you need a solid sample of each silver halide free from any other chemical;
- silver halides can darken on exposure to light;
- the results of your investigation need to be supported by measurements;
- what new chemicals are made as a result of any darkening.



Wear eye protection.

Check your plan with your teacher before carrying out your investigation.

Silver halides: Fact file
• they are insoluble in water
• made by a process which involves the mixing of ions, called ionic precipitation
• some are light-sensitive compounds
• darkening is due to the formation of metallic silver

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Making shadow prints

In this activity you are going to use light-sensitive photographic paper to make a picture of the shadow cast by an object, or a group of objects.

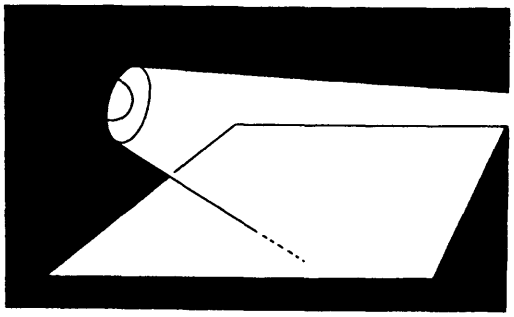
The photographic paper is coated with silver halides which react to form metallic silver when exposed to white light.

Even a small amount of light will 'fog' the paper, so you must take care to keep it in the dark until you are ready to use it.

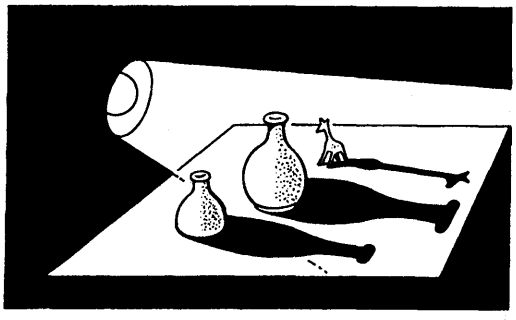
It is safe to use red light while you develop and fix the image you have made.

◆ *Taking the shadow photograph*

- Work in a darkened room.



Set up a lamp to shine on a piece of white paper.



Arrange objects on the white paper to give interesting shadows.

- Switch off the lamp and replace the sheet of white paper with photographic paper.
- Turn on the light. An exposure time of between 5 and 10 seconds should give a good result, though you might wish to investigate the effect of using a longer or shorter exposure time.
- Quickly place your unprocessed shadow print between the pages of a book to keep it in the dark.

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Making shadow prints

◆ *Processing the shadow print*

Processing is a two-stage operation:

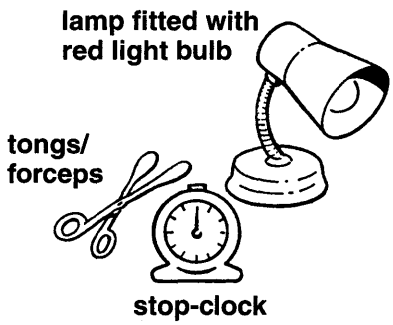
- first the image must be developed
- then fixed to make it permanent.

Your teacher will tell you how long to leave the print in the developer and fixer baths and how to use the chemicals safely. Take care because many photographic chemicals are hazardous.

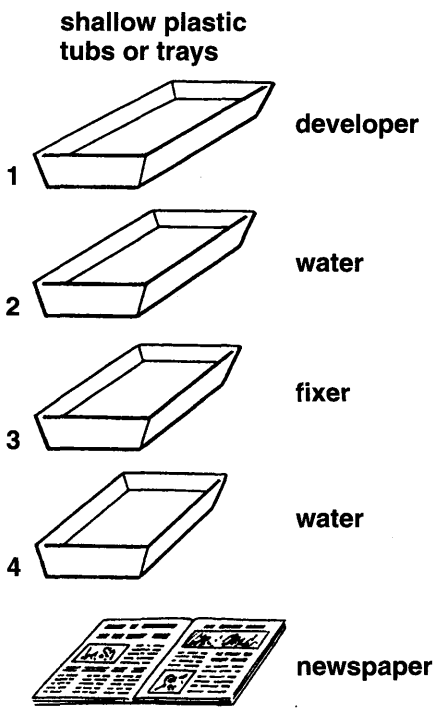
The apparatus for processing the shadow prints is shown below.

- Arrange the trays in order, so that it is easy to move your picture from one to the next.
- While your picture is in each of the baths, rock the tray gently so that the liquid is kept moving over the paper.

You will need a lamp fitted with a red light bulb, tongs or forceps and a stop-clock.



- Develop your print
- Wash it very quickly
- Put it in the fixer bath
- Wash it quickly again
- Put it on newspaper to dry



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Now you see it . . .





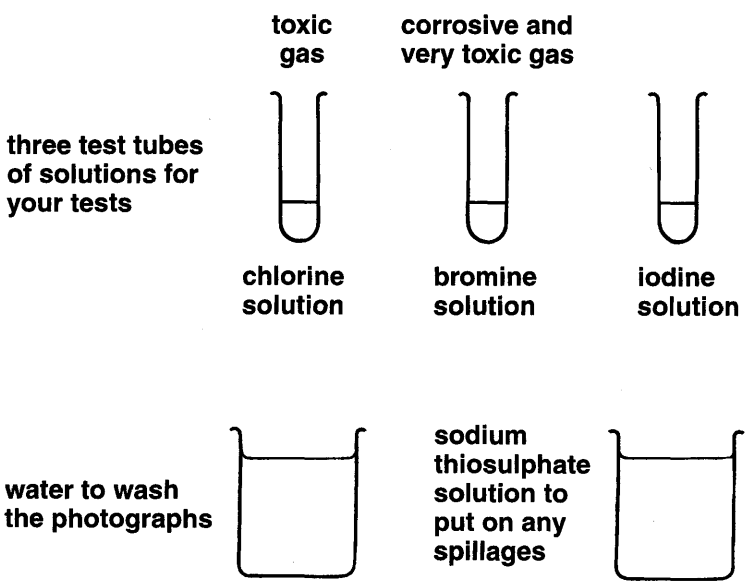
. . . Now you don't!

In this series of experiments you are going to find out about some of the properties of chlorine, bromine and iodine when these elements are dissolved in water.

◆ *Removing silver from black and white photographs*

- Imagine that you are a photographer and that you want to remove an object from a black and white print. Perhaps its presence spoils the composition of the picture. Which halogen solution will you use?

In the fume cupboard you will find the following things:





Wear eye protection.

Use the halogen solutions in a fume cupboard.

Do not remove photographs from the fume cupboard until they have been washed.

Do not spill solutions on your skin. If you do, wash the affected area thoroughly and ask your teacher for help.

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Now you see it . . .

- Cut three strips from a black and white photograph. Each strip should be about 1 cm wide.
- Holding one strip with forceps or tongs, dip it into the chlorine solution for 30 seconds, then remove it and wash it in a beaker of water.
- Repeat with the second strip, dipping it first into bromine solution and then washing it in water.
- Repeat with the third strip, using iodine solution this time.

Q1 Which solution would you use to remove part of a black and white photograph? Explain your choice.

- You might try carefully to remove an object from a photograph using your chosen halogen solution and a paintbrush.

◆ *Bleaching the dyes from colour photographs*

When a colour photograph is processed, all the silver is removed and only a mixture of coloured dyes remains. Could a photographer use a solution of chlorine, bromine or iodine to remove or lighten a coloured image?

- Cut three strips from a colour photograph. Each strip should be about 1 cm wide.
- Dip one strip into the chlorine solution for 30 seconds, then wash it, as before.
- Repeat using the bromine and iodine solutions, using the second and third strips of colour photograph.

Q2 Which solutions affect or bleach the dyes and which do not?

Q3 When the photograph has been bleached, which colours have been affected most and which least?

◆ *Using solutions of halogens to bleach other dyes*

- Investigate the bleaching effect of the halogen solutions on other coloured materials. You might try dipping in blue litmus paper or adding a drop of blue ink to each solution.

Q4 Which of the solutions bleach the other dyes you used?

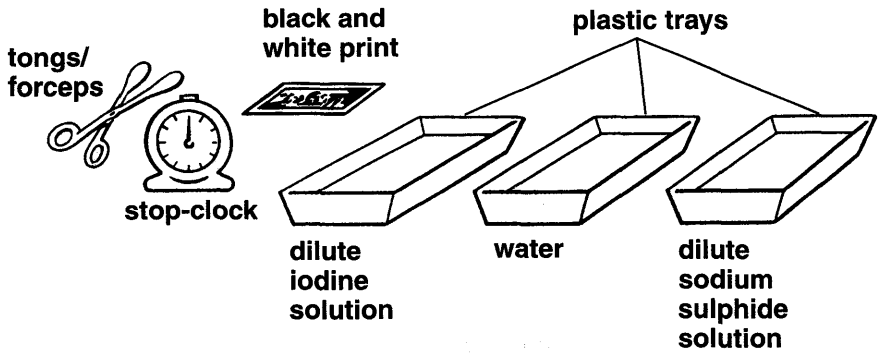
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Now you see it . . .

◆ *Using iodine solution to make 'sepia' prints*

It was once fashionable to turn black and white tones into much softer brownish tones. This kind of photograph became known as a 'sepia' print. You can do this using dilute iodine solution.

- Set up this apparatus in a fume cupboard.





Wear eye protection.

- Place the black and white print into the dilute iodine solution until the image is faded and yellow.
- Wash the faded print in the tray of water.
- Place the faded print into the sodium sulphide solution (**care – corrosive**) until the image darkens.
- Wash your finished 'sepia' print and allow it to dry.

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Close-up on the halogens

This sheet gives some information about the elements chlorine, bromine and iodine, and their compounds. Use it to make predictions about the properties of two other halogens, fluorine and astatine, and their compounds. Write your predictions in the spaces in the tables below.

◆ *What do they look like?*

element	state
fluorine	
chlorine	gas
bromine	liquid
iodine	solid
astatine	

silver halide	how it looks
silver fluoride	
silver chloride	white solid
silver bromide	cream solid
silver iodide	pale yellow solid
silver astatide	

◆ *What are their chemical reactions?*

	with coloured dyes	with reactive metals to form salts	with the non-metal, hydrogen	reactivity of silver halides with light
chlorine	removes blues and greens	reacts very vigorously	reacts explosively if light present	reacts quickly
bromine	removes blues and greens		reacts when heated with a catalyst	
iodine	no bleaching action	reacts quite quickly	reacts partially when heated with a catalyst	reacts very slowly

Q1 How does the reactivity of the halogen vary from chlorine to iodine?

Q2 How would you expect the reactivity of fluorine to compare with that of chlorine?

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Close-up on the halogens

◆ *Chemical formulas*

of the atoms	of the elements	of the sodium halide salts
F		
Cl	Cl_2	NaCl
Br	Br_2	NaBr
I	I_2	NaI
At		

Q3 Fill in the formulas you would expect in the empty boxes.

◆ *Atomic structure and bonding*

	total no. of electrons	no. of outer electrons	no. of electrons needed to complete outer shell
fluorine	9	7	
chlorine	17	7	1
bromine	35	7	1
iodine	53	7	1
astatine	85	7	

- Draw 'dot and cross' diagrams to represent the bonding in:

  (a) a molecule of chlorine, Cl_2,

  (b) a molecule of iodine, I_2.

Q4 How are the atoms of chlorine and iodine achieving a full outer shell of electrons?

- Draw 'dot and cross' diagrams to represent the bonding in

  (a) sodium fluoride, NaF,

  (b) sodium iodide, NaI.

Q5 How are the atoms of fluorine and iodine achieving a full outer shell of electrons?

Q6 Explain why there is a family resemblance between the halogen elements.

- Make up your own 'Fact File' on the halogens, similar to that used earlier in this section for silver halides (*Student Sheet 3.2*, page 57).

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A host of uses

How many uses of chlorine, bromine and iodine can you find by reading these cuttings? You might record your findings in a table, and add those uses you have learned about earlier in the lesson.

Halothane
General inhalation anaesthetic.
2-Bromo-2-chloro-1,1,1-trifluoroethane

GAS! GAS! Quick, boys! An ecstacy of fumbling,
Fitting the clumsy helmets just in time;
But someone was yelling out and stumbling
And floundering like a man in fire or lime ...
Dim, through the misty panes and thick green light
As under a green sea, I saw him drowining.

from a poem by Wilfred Owen

Antiknock agents, e.g. tetramethyl lead break down to form lead oxides dur[illegible] combustion. If these are allowed to build up, engine performance will deteriorate. 1,2-dibromoethane converts these lead oxides into lead halides which are volatile at engine operating temperatures and are removed in the exhaust gases.

Organochlorine insecticides. the first synthetic insecticides and introduced during World War II. DDT discovered by Paul Mueller in Switzerland in 1939 and was first produced in 1942 in the USA. The Nobel Prize was awarded in 1948 for its discovery. Within 8 years it had saved 5 million lives and hel[illegible] over 100 million illnesses

THE CONCENTRATION of halons in the atmosphere over Antarctica and Tasmania is rising by about 25 per cent a year, according to an Australian scientist. Halons, like chlorofluorocarbons (CFCs), destroy the ozone layer. These are widely used in fire extinguishers. The most widely used halons (bromochlorodifluoromethane), 1301 (bromotrifluoromethane), to 10 Treated water from the contact tank in destructive on the m Clevedon Reservoir for distribution.

The **quality** of the water is always being checked are numerous systems to ensure that a good source maintained.
Should a high or low chlorine residual chlorinator and well pumpsets will [illegible]
and warnings will be give[illegible]

A safety expert from a major industrial concern told Jerusalem Post that the bromine had been packed according to international safety standards but that yesterday's accident highlighted the need for more careful monitoring of the transportation of dangerous chemicals. It was the first accident ever in Israel involving bromine, a deep rust-coloured liquid used in industry, agriculture and medicine.

Poly(chloroethene) or polyvinyl chloride is obtained when chloroethene (vinyl chloride) is polymerised. The presence of chlori[illegible] atoms makes this plastic a tough, rigid material can be used to make pipes and guttering.

Iodine is taken into the body either as iodide ions, added in small quantities to ordinary salt and also present in very small quantities in drinking water, or as iodate(V) ions, present in marine plants and animals. Iodide ions are also present in cows' milk, particularly in winter when cows are given iodised salt licks to supplement their diet.

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Piecing together the jigsaw

2 lessons

In this section students have to piece together information about many of the substances they have encountered throughout this unit and the course as a whole.

They compile a list of the properties associated with different types of structures: giant ionic, giant metallic, giant covalent and simple covalent. They then identify the structures present in some substances, given their properties.

Students consider patterns in the arrangement of elements in the Periodic Table. The link between the electron arrangement of an atom and its position in the Periodic Table is reinforced. The opportunity is given for students to remind themselves of the content of other units.

ACTIVITY	OUTLINE	KEY POINTS
Discussion 1	Students discuss how substances encountered in this unit can be usefully classified in terms of their structure and bonding.	Pure substances can be classified in terms of the bonds they contain and the ways in which their component particles are arranged.
Classifying data 1 (Student Sheets 4.1 and 4.2)	Students use books to find out about the structure, bonding and properties of some of the substances. They draw up a list of properties associated with each different type of structure. Finally, they use the Periodic Table to predict which pairs of elements will combine ionically and which will combine covalently.	Substances with low melting points usually contain simple molecules. There are weak attractive forces between the molecules, but strong covalent bonds holding the atoms together within the molecules. Compounds which conduct electricity when molten or in solution contain ions. The ions are held in place in the solid by ionic bonds between positive and negative ions.
Practical work (Student Sheet 4.3)	Students investigate the properties of some solids, and use a key to obtain information about their structure and bonding.	Substances with high melting points which are insoluble in water and do not conduct electricity when molten usually contain giant structures in which the atoms are covalently bonded. Substances with high melting points which conduct electricity when solid usually contain giant structures held together by metallic bonds.
Homework suggestion (Student Sheet 4.4)	Students identify the structures of some substances from their properties.	
Discussion 2	The link between the Periodic Table and electron arrangements in the atoms is reinforced and the link is made with combining power.	The number of outer electrons for any atom is related to their group in the Periodic Table. The combining power of atoms depends on the number of outer electrons.
Classifying data 2 (Student Sheets 4.5 and 4.6)	Students use data books and textbooks to complete tables linking the position of elements in the Periodic Table to their electron arrangement and properties.	Typical properties of noble gases and transition metals.

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ACTIVITY	REQUIREMENTS FOR THE ACTIVITY
Discussion 1	
Classifying data 1	Student Sheet 4.1 *Properties of substances* Student Sheet 4.2 *The Periodic Table* chemistry data books and textbooks
Practical work	Student Sheet 4.3 *Using a key to identify structures* test tubes, test tube holders, test tube racks 100 cm^3 beakers, spatulas, Bunsen burners, glass rods connecting leads with crocodile clips, torch bulbs in holders battery or power pack (6 V), two carbon electrodes salt, labelled 1 sand, labelled 2 grated paraffin wax, labelled 3
Homework suggestion	Student Sheet 4.4 *Identifying the structures of compounds*
Classifying data 2	Student Sheet 4.5 *The Periodic Table and electron arrangements* Student Sheet 4.6 *The Periodic Table and combining power* chemistry data books and textbooks

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Piecing together the jigsaw

Students collect information about elements and compounds and see that the Periodic Table can be used to classify this information. The link between the position of an element in the table, its electronic configuration and its properties is established.

Discussion 1: *Substances mentioned in this unit*

You could compile, with the help of the students, a list of substances that have been mentioned in this unit. The list might include the following:

aluminium
aluminium oxide
ammonia
bromine
copper sulphate
carbon dioxide
carbon monoxide
chlorine
diamond
graphite
hexane
iodine
lithium
lithium oxide
magnesium
magnesium oxide
methane
nitrogen
nitrogen monoxide
octane
oxygen
potassium bromide
potassium chloride
potassium iodide
sodium carbonate
sodium chloride
sodium hydrogencarbonate
sodium sulphate
water
zinc

It would be wise to omit the silver halides because the purpose of the activity which follows is to classify substances in terms of their structure and bonding, and these particular compounds are partly ionic, partly covalent.

Ask the students to suggest ways in which the substances could be classified. They might mention the following:

- metallic/non-metallic;
- state at room temperature;
- type of particle present;
- type of bond present;
- arrangement of particles in the substance.

The physical properties of a substance depend largely on the way in which its component particles are arranged and what type of bond is involved. So the classification of substances in terms of their structure and bonding is very useful.

Classifying data 1: *Properties of substances*

Student Sheets 4.1 and 4.2 accompany this activity.

The students could work collaboratively to find information about the structure, bonding and properties of the substances they have listed. They might seek this information from data books and textbooks, and from their own notes.

Student Sheet 4.1 is a pro-forma which will help to focus the students' attention on information which is relevant to the exercise. The class could then come together and, with the help of the teacher, the substances could be grouped together according to the types of structure and bonding they possess.

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This classification divides the substances into four groups.

1 **SIMPLE COVALENT MOLECULES**

e.g. ammonia, bromine, carbon dioxide, carbon monoxide, chlorine, hexane, iodine, methane, nitrogen, nitrogen monoxide, octane, oxygen, water.

2 **GIANT IONIC STRUCTURES**

e.g. aluminium oxide, copper sulphate, lithium oxide, magnesium oxide, potassium bromide, potassium chloride, potassium iodide, sodium carbonate, sodium chloride, sodium hydrogencarbonate, sodium sulphate.

3 **GIANT COVALENT STRUCTURES**

e.g. diamond, graphite.

4 **GIANT METALLIC STRUCTURES**

e.g. aluminium, lithium, magnesium, zinc.

The students might then work together in groups, each group compiling a list of general properties associated with one type of structure. You might guide the students by asking that they keep the following questions in mind whilst they are working.

For substances with a particular type of structure:

- are melting and boiling points high or low;
- are there weak or strong forces holding the particles together;
- do they conduct electricity;
- are they soluble in water?

Make sure that the finished lists cover the key points, and that the students have a written note of each one.

In the second part of the activity, the students shade in the Periodic Table, given on *Student Sheet 4.2*, the elements present in the substances they have just classified.

The purpose of the exercise is for the students to decide which pairs of elements are likely to bond together ionically, and which covalently. If they use their knowledge of the substances they have just classified, they should be able to work out what type of bond will be formed when other pairs of elements combine.

Some of the ionic compounds they have met in the unit are aluminium oxide, lithium oxide, lithium chloride, magnesium oxide and sodium and potassium salts. They should be able to work out that ionic bonds are formed when a metallic element combines with a non-metallic one.

Some of the covalent compounds they have met are carbon dioxide, methane, water, ammonia and hexane. Covalent bonds are formed when non-metallic elements combine. You might ask the students to suggest which type of bond exists within elements like oxygen, chlorine, bromine and iodine. They should be able to work out, from the properties of the elements, that these, too, contain covalent bonds.

Finally, ask the students to predict the type of bonds present in other substances, such as calcium oxide and sulphur dioxide.

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Practical work: *Using a key to identify different structures*

Student Sheet 4.3 accompanies this activity.

The students investigate the properties of three solids and then classify each one in terms of its structure, by means of a key. One of the solids should be a simple molecular substance, one a giant covalent substance and the third a giant ionic substance. Suitable solids are finely grated paraffin wax, sand and salt. The key is, of course, a generalisation and the students should be asked to name some exceptions; e.g. calcium carbonate is an insoluble ionic compound and sugar is a soluble simple molecular compound.

Homework suggestion: *Identifying the structures of compounds*

Student Sheet 4.4 accompanies this activity.

The students answer questions about the structure of various different substances, given their properties.

Answers to questions on *Student Sheet 4.4*

Q1 Potassium chloride has a giant ionic structure.

Q2 Carbon monoxide forms simple molecules with covalent bonding.

Q3 Sand has a giant covalent structure.

Q4 Sugar forms simple molecules with covalent bonding.

Q5 Rust has a giant ionic structure.

Discussion 2: *Patterns and the Periodic Table*

This is an ideal opportunity to draw together ideas developed on the Periodic Table and electron arrangements. Following the discussion there are a number of tables for the students to complete, with or without the help of the teacher, to consolidate the links made during the discussion. The idea of combining power, which was covered in FACT FILE 7: *Transporting Chemicals*, can now be developed further. From discussion the following patterns can be emphasised:

- Group number is the same as number of electrons in the outer shell (except for transition metals and noble gases).
- Elements in the same group have the same number of electrons in the outer shell.
- The position of an element in a group is linked to its reactivity.

  Alkali metals and alkaline earth metals are more reactive as you go down the group.

  Halogens are more reactive as you go up the group.
- The combining power of an element is linked to number of electrons in the outer shell.

  Ionic compounds: combining power is the same as the number of charges that an ion has.

  Covalent compounds: combining power is the same as the number of electrons an atom of an element shares with another atom.
- The noble gases have complete outer shells and are, therefore, unreactive.

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Classifying data 2: *Electron arrangements and combining power*

Student Sheets 4.5 and 4.6 accompany this activity.

Students complete information tables which are presented in a structured manner so as to consolidate information, ideas and concepts which have previously been developed throughout the course. With this unit coming late in the course, this activity is ideal revision for the students whilst at the same time it relies on good data-handling skills. If students are confident with databases then they could make their own database instead of completing the Student Sheets.

The amount of help given depends on the nature and ability of the teaching group. The Student Sheets could be given as homework.

Answers to questions on *Student Sheet 4.5*

Q1 (a) 1 electron

(b) 2 electrons

(c) 7 electrons

Q2 Group I - by losing one electron to form ions with a charge of 1+.

Group II - by losing two electrons to form ions with a charge of 2+.

Q3 They each tend to lose the single electron from their outer shell. This is more easily done from atoms with larger atomic radii as the electron is further from the positively charged nucleus. Therefore the elements become more reactive as you move down the group.

Q4 Group VII atoms gain one electron to form ions with a charge of 1–.

Q5 Atoms with small atomic radii gain electrons more readily than larger atoms as the electron gained is closer to the positively charged nucleus. This is why the reactivity of halogens increases as you go up the group.

Answers to questions on *Student Sheet 4.6*

Q1 The charge is equal to the combining power.

Q2 The charge is equal to the combining power.

Q3 The combining power is equal to the number of covalent bonds formed.

Q4 (a) The Haber process for the production of ammonia uses iron as a catalyst.

(b) The hydrogenation of fats to make margarine uses nickel as a catalyst.

(c) The Contact process for the production of sulphuric acid uses vanadium(V) oxide as a catalyst in the reaction between sulphur dioxide and oxygen to form sulphur trioxide.

Q5 The noble gases have complete outer shells of electrons so their electron arrangement is stable and they are unreactive.

Q6 Helium has a low density and it is unreactive and so, unlike hydrogen, which could be used, it is not flammable and is therefore safer.

Q7 An argon atmosphere would be inert and so it would avoid oxidation of the hot metal.

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Properties of substances

Name of substance		Formula of substance
State at room temperature	Melting point (°C)	Boiling point (°C)
Appearance		
Does the substance dissolve in water?		
Does the substance conduct electricity (a) when solid?	(b) when molten?	(c) when in solution?
Does the substance burn easily?		
What type of particle is present?		
What type of structure does the substance form?		

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The Periodic Table

The Periodic Table of the Elements

	Group I	Group II										Group III	Group IV	Group V	Group VI	Group VII	Group 0
1								1 **H** Hydrogen 1									4 **He** Helium 2
2	7 **Li** Lithium 3	9 **Be** Beryllium 4										11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
3	23 **Na** Sodium 11	24 **Mg** Magnesium 12										27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulphur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
4	39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	64 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
5	85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	99 **Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
6	133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57 *	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	**Po** Polonium 84	**At** Astatine 85	**Rn** Radon 86
7	**Fr** Francium 87	226 **Ra** Radium 88	227* **Ac** Actinium 89 †															

*	140 **Ce** Cerium 58	141 **Pr** Praseodymium 59	144 **Nd** Neodymium 60	**Pm** Promethium 61	150 **Sm** Samarium 62	152 **Eu** Europium 63	157 **Gd** Gadolinium 64	159 **Tb** Terbium 65	162 **Dy** Dysprosium 66	165 **Ho** Holmium 67	167 **Er** Erbium 68	169 **Tm** Thulium 69	173 **Yb** Ytterbium 70	175 **Lu** Lutetium 71
†	232 **Th** Thorium 90	231 **Pa** Protactinium 91	238 **U** Uranium 92	237 **Np** Neptunium 93	**Pu** Plutonium 94	**Am** Americium 95	**Cm** Curium 96	**Bk** Berkelium 97	**Cf** Californium 98	**Es** Einsteinium 99	**Fm** Fermium 100	**Md** Mendelevium 101	**No** Nobelium 102	**Lr** Lawrencium 103

key

a **X** b

a relative atomic mass

X atomic symbol

b atomic number – order of element in the Periodic Table

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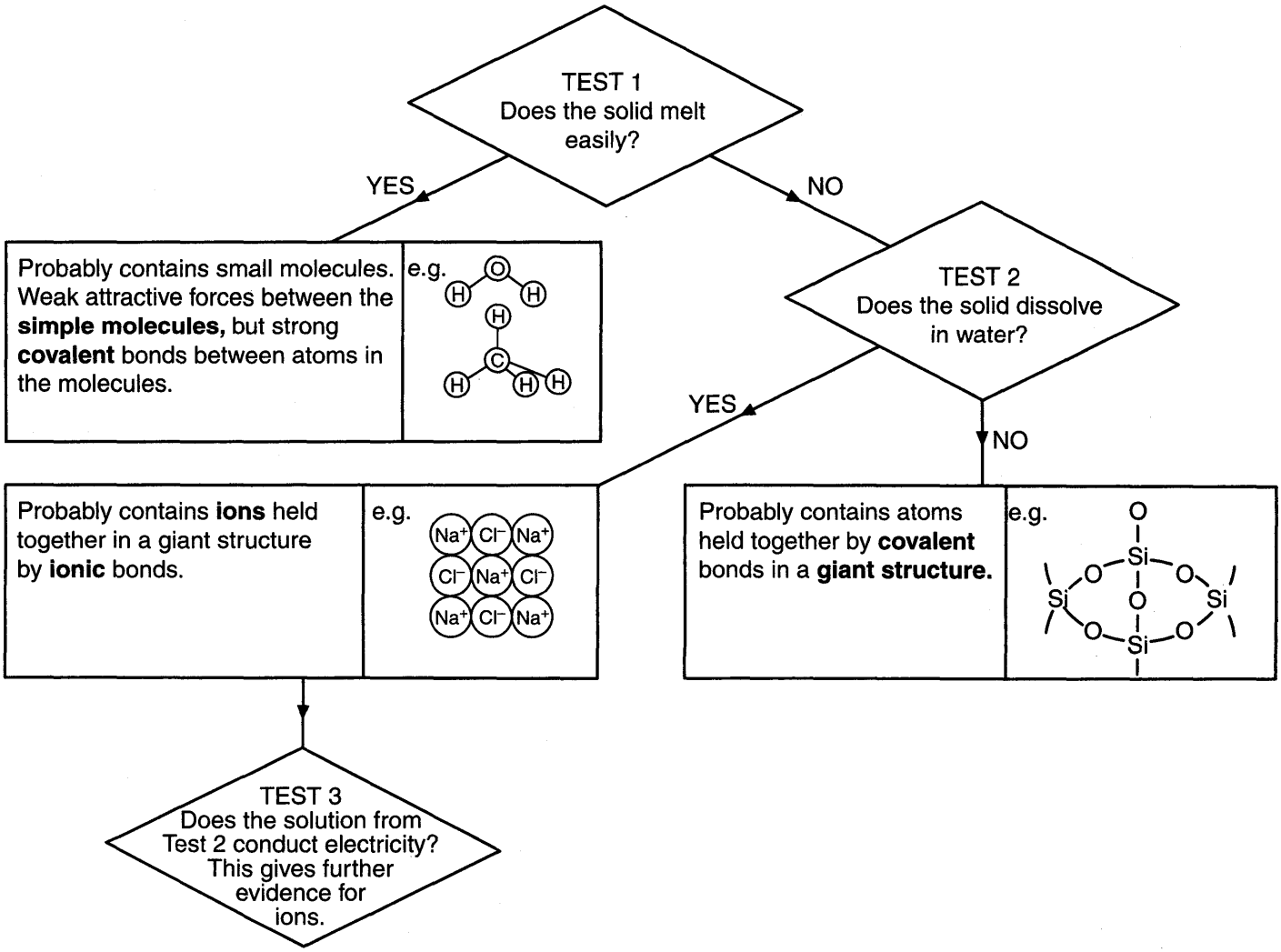
Using a key to identify different structures

- In this experiment you are going to investigate the properties of solids (labelled A, B and C) and use a key to obtain information about their structure and bonding.
- You need to carry out three tests on each substance:

  Test 1: heating, for information on their melting points;

  Test 2: adding water, for information on their solubility;

  Test 3: testing for conduction of electricity.
- Study the key.
- Carry out the three tests on each of the three solids.
- Classify the solids as simple molecular, giant covalent or ionic.



Some ionic solids are not soluble in water. The best way to check whether or not an insoluble substance contains ions is to melt it, and see if the molten substance conducts electricity. If it does, the substance contains ions. Unfortunately, the temperatures needed are higher than can be achieved using a Bunsen burner, so this part of the key has been omitted.

- Decide where you would put this test, if you wanted to add it to the key.
- Adapt the key so that it could be used to identify substances with giant metallic structures.

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Identifying the structures of compounds

Different compounds have different properties.

The properties of a compound are determined by its structure and by the bonds between particles in it.

◆ *A table of properties*

The table below summarises the properties of compounds with different types of structure.

type of structure	properties		
	melting point	**conduction of electricity when molten**	**solubility in water**
giant structures of atoms with covalent bonds between atoms	high	poor	insoluble
giant structures of ions held by ionic bonds	high	good	usually soluble
simple molecules with covalent bonds between atoms in the molecules	low	poor	some are soluble, others not

- Use the table above to predict the structures of the following compounds.

Q1 Potassium chloride dissolves in water, melts at 808°C and conducts electricity when molten.

Q2 Carbon monoxide does not dissolve in water and melts at –205°C.

Q3 Sand does not dissolve in water, melts at 1610°C and is a poor conductor of electricity when molten.

Q4 Sugar (sucrose) dissolves in water, melts at 186°C but does not conduct electricity when molten.

Q5 Iron oxide (the main component of rust) does not dissolve in water, melts at 1565°C and conducts electricity when molten.

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The Periodic Table and electron arrangements

- Complete the following table and answer the questions.

element	electron arrangement	number of electrons in outer shell	group number
Hydrogen	1	1	–
Helium	2	2	0
Lithium	2,1	1	1
Beryllium			
Boron			3
Carbon			
Nitrogen			5
Oxygen			
Fluorine			
Neon	2,8	8	0
Sodium			
Magnesium			
Aluminium			
Silicon			4
Phosphorus			5
Sulphur			
Chlorine			
Argon	2,8,8	8	0
Potassium			
Calcium			

Q1 What do you notice about the number of electrons in the outer shell of:

(a) Group I elements;

(b) Group II elements;

(c) Group VII elements?

Q2 How do the metals in Groups I and II form ions?

Q3 The reactivity of Group I and II metals increase as you go **down** the group. Explain this trend by considering the electron arrangement of the metals in each group and the increase in atomic radii down the group.

Q4 How do the halogens in Group VII form ions?

Q5 The reactivity of Group VII non-metals towards metals increases as you go **up** the group. Explain this trend by considering the electron arrangement of the halogens (Group VII) and the increase in atomic radii down the group.

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The Periodic Table and combining power

- Complete the following tables and answer the questions.

◆ *Metals*

element	number of electrons in outer shell	charge on ion	combining power
Lithium	1	1+	1
Sodium			
Potassium			
Magnesium			
Calcium			
Aluminium			

Q1 What do you notice about the charge on the ion and combining power for the metal elements?

◆ ***Non-metals:*** when reacting with metals

element	number of electrons in outer shell	charge on ion	combining power
Oxygen			
Fluorine			
Sulphur	6	2–	2
Chlorine			
Bromine			
Iodine			

Q2 What do you notice about the charge on the ion and combining power of the non-metal elements when reacting with a metal?

◆ ***Non-metals:*** when reacting with other non-metals

element	number of electrons in outer shell	number of covalent bonds formed by one atom of the element	combining power
Hydrogen	1	1	1
Carbon			
Nitrogen	5	3	3
Oxygen			
Fluorine			
Phosphorus			3
Sulphur			
Chlorine			

Q3 What do you notice about the number of covalent bonds formed and the combining power of the element when reacting with a non-metal?

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The Periodic Table and combining power

◆ *Transition metals*

element	charge on ions	compounds white or coloured?	special properties	combining powers
Iron	+2 and +3		Can act as catalyst	
Copper	+1 and +2		Can act as catalyst	

Q4 Which industrial process uses:

(a) iron as a catalyst;

(b) nickel as a catalyst;

(c) vanadium(V) oxide as a catalyst?

◆ *The noble gases*

element	symbol for the element	uses	outer shell of electrons
Helium		weather balloons 'breathing mixture' for divers	full
Neon		neon lights	full
Argon		light bulbs welding	full
Krypton		powerful lamps	full
Xenon		powerful lamps	full
Radon		treatment for tumours	full

Q5 The noble gases are generally inert (unreactive). Explain this by considering the electron arrangement of the gases.

Q6 Give **two** reasons why helium is a good choice for weather balloons.

Q7 What would be the advantage of welding a metal in an atmosphere of argon instead of air?

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Different kinds of compounds

The atoms of non-metal elements have nearly full outer layers of electrons. When non-metal elements join together they share electrons, forming covalent bonds and so completing their outer layers.

Atoms of metal elements have only 1, 2 or 3 outer electrons – they must react in a different way.

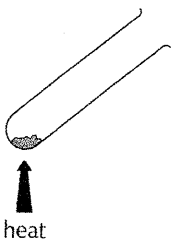
You will be given some compounds containing only non-metals, and some which have metal *and* non-metal atoms in them. You can test some of the differences in their properties.



Wear eye protection.

Melting point

Q1 Which substances are liquids at room temperature and which are solids?





Avoid making and inhaling fumes.

Take each solid substance in turn. Put a small sample in an ignition tube (a very small test-tube) and heat it gently at first, then more strongly. Stop heating if the sample melts.

Q2 Which substances melted?

Solubility in water

For each substance in turn, shake a very small sample with water in a test-tube.

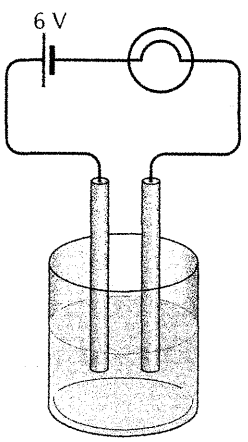
Q3 Which samples dissolve in water?

Electrical conductivity

Use this apparatus to test

a Sugar solution (sugar is a covalent compound, containing carbon, hydrogen and oxygen)

b Salt solution (salt is a compound of sodium with chlorine)





Wear eye protection.



Do not inhale the gas produced. Tell your teacher if you suffer from asthma.

The bulb in the circuit acts as an indicator to tell you whether any electric current is passing.

Q4 Which solution conducts electricity?

Now repeat this experiment, using copper chloride solution.

Look very carefully at the two carbon rods while the electricity is switched on.

Q5 What do you see at the cathode (the rod connected to the negative terminal)?

Q6 What substance is being extracted from the solution?

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Group I and Group II metals

With your group, look up information about these metals and the compounds they form. You can also carry out some experiments. Use the information you collect to fill in tables of information about the metals.

Try to find out

- which group of the Periodic Table each metal belongs to (e.g. sodium is in Group I).
- the arrangement of electrons in its atoms (e.g. sodium has the arrangement 2, 8, 1).
- the formula of the oxide of each metal (e.g. sodium oxide is Na_2O).
- the formula of the hydroxide, and whether it is acidic or alkaline in solution.
- the formula of the chloride of each metal (e.g. sodium chloride is NaCl).
- the formula of the carbonate of each metal, and the effects of heat or acids on the carbonates.



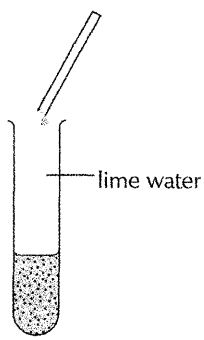
Wear eye protection.

Experiment 1: Hydroxides of these metals

You have been given solutions of the hydroxides of each of the metals. Test each solution with universal indicator to find out whether it is acidic, neutral or alkaline.

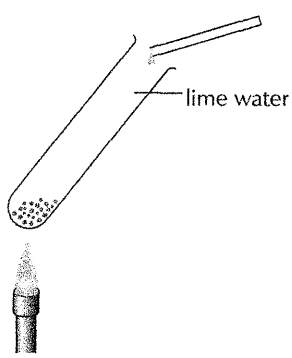
Experiment 2: The action of acid on carbonates of these metals

- You have been given solid samples of the carbonate of each metal.
- Take a small amount of the sample in a test tube.
- Add enough dilute hydrochloric acid to cover the sample.
- Pick up a drop of lime water on the end of a glass rod and hold it inside the mouth of the test-tube.
- What gas is given off?



Experiment 3: The effect of heat on the carbonates of these metals

- Put enough sample in a test-tube to fill the rounded tip.
- Hold the tube at an angle and heat gently.
- Hold a drop of lime water in the mouth of the tube.
- If nothing happens, heat more strongly.



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Some properties of Group I and Group II Metals and their compounds

metal	group	electron arrangement	formula of oxide	formula of hydroxide	acid/alkaline	formula of chloride	formula of carbonate	effect of acid	effect of heat
lithium									
sodium	I	2, 8, 1	Na_2O	NaOH	alkaline	NaCl	Na_2CO_3	carbon dioxide evolved	no change
potassium									
magnesium									
calcium									

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Compounds with chlorine, bromine and iodine

The elements in Group VII of the Periodic Table are called **halogens**. Chlorine, bromine and iodine are the best known members of this family.

You can test for simple salts of these elements using silver nitrate solution.

Making a sample of the silver halides

The silver part is provided by using **silver nitrate solution**, this contains **silver ions**.

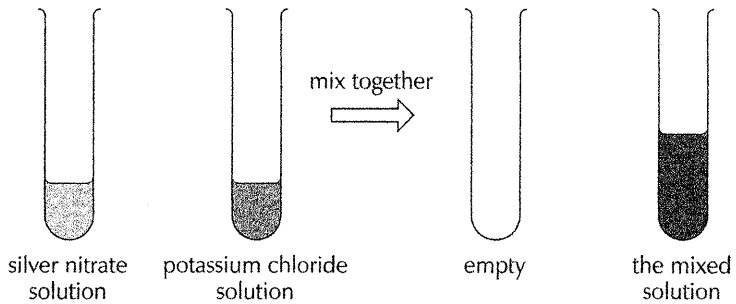


Wear eye protection.

The halide part is provided by using:

- **potassium chloride solution** to make silver chloride
- **potassium bromide solution** to make silver bromide
- **potassium iodide solution** to make silver iodide.

When made this way the silver halides appear as solids.



Q1 What do you see in each tube?

Keep the test tubes with the mixed solutions. Stand them in a test tube rack on a window sill where they get plenty of light.

Comparing the activity of the halogens

When you were comparing the reactivity of different metals, you found that a more reactive metal could **displace** a less reactive one from a compound (e.g. iron or magnesium displace copper from copper sulphate).



Take great care with chlorine water and bromine water.

Do not inhale the fumes.

Avoid skin contact with all of these chemicals.

In the same way, a more reactive halogen can displace a less reactive one.

- Stand three test tubes in a rack.
- In one put 1 cm depth of potassium chloride solution.
- In the next, put 1 cm depth of potassium bromide solution.
- in the next, put 1 cm depth of potassium iodide solution.
- Add 2–3 drops of chlorine water (**CARE**) to each test-tube.

Q2 What do you see in each tube?

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Compounds with chlorine, bromine and iodine



Wear eye protection.

- Wash out the tubes, then refill them as before.
- This time, add 2–3 drops of bromine water (**CARE**) to each tube.

Q3 What do you see in each tube?

- Wash out the tubes and refill them as before.
- This time, add 2–3 drops of iodine solution to each tube.

Q4 What do you see in each tube?

Q5 Explain how your results help you to decide which is the most reactive element.

Q6 Write chemical equations for each of the reactions which happened.

The effect of light on silver halides

Look carefully at the tubes with silver halides in.

Q7 What has happened to the contents of the tubes?

Some extra information

A sample of silver chloride, silver bromide or silver iodide (silver halides) will get darker in colour if it is exposed to light. Light energy causes a reaction which produces silver metal.

Photographic film has a layer of silver halide spread over a transparent material like cellulose acetate.

When light falls onto a silver halide crystal a tiny part of that crystal may react to form metallic silver. You can't see any dark colour, but a record of the scene, a latent image, has been formed. When the film is processed the developer turns much more of the crystal to silver and a full image develops.

If you expose a film to too much light, all the silver halide crystals will begin to react to form silver. When you develop the film it will be black all over.

So film must be kept away from light; many modern films are packed in plastic cassettes so that you can load and unload your camera in the light.

After developing, the image is **fixed** by washing it in a solution which dissolves away the unreacted silver halide, so that no more of the crystals can react to form silver. These chemicals are called **fixers**.

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Be a transition metal detective

Transition metals often form coloured compounds. These colours can be used to identify the metals which are present.

You will be shown labelled bottles containing salts of different transition metals. Some sodium chloride solution is also on show, so that you can compare a group I metal salt to the transition metal salts.

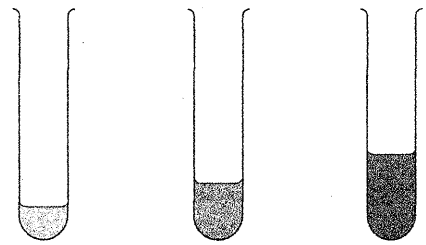
Q1 Write down the colour which is typical of each metal.

The sodium hydroxide test

Test each of the labelled samples in turn (including the salt solution).

Place about 1 cm depth of the sample solution in a test tube.

Add an equal volume of sodium hydroxide solution, and shake the tube gently from side to side to mix the liquids.



Q2 Write down what you see for each of the samples.

Now add about 2 cm depth more sodium hydroxide solution. Mix the liquid in the tube by gently stirring with a glass rod.

Q3 Write down what you see when extra sodium hydroxide is added.



Wear eye protection.

The ammonium hydroxide test

Test each of the labelled samples in turn (including the salt solution).

Place about 1 cm depth of the sample solution in a test tube.

Add an equal volume of ammonium hydroxide solution, and shake the tube gently from side to side to mix the liquids.

Q4 Write down what you see for each of the samples.

Now add about 2 cm depth more ammonium hydroxide solution. Mix the liquid in the tube by gently stirring with a glass rod.

Q5 Write down what you see when extra ammonium hydroxide is added.

Identifying unknown samples

You will be given some samples which are identified only by numbers or letters. Test each one to see which metal is present in the dissolved substance.

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science

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