Zig Zag Education

Chemistry

GCSE (9–1) | AQA | 8464

How to Pass

GCSE AQA Combined Science: Chemistry Paper 1

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Teacher's Introduction

The aim of this resource is to provide your students with revision materials, guidance and practice to help them secure a grade 5 in **GCSE (9–1) AQA Combined Science (Trilogy) Chemistry Paper 1**.

Remember!
Always check the exam board website for new information, including changes to the specification and sample assessment material.

As teachers, we always want our students to attain the highest grades possible. For many of your students entered for the Foundation Tier paper, the challenge is to secure a grade 5. With the low grade boundaries on the test papers, the focus of this resource is on some key and easily accessed areas of the specification for the Chemistry Paper 1 topics. The 'quick check' questions are similar to the questions that appear on the exam paper; the worked examples will guide you through how to answer such questions.

The specification for topics in Paper 1 has been divided into five chapters. Each chapter consists of the following:

- A student-focused background section, explaining and summarising the main points of the specification and identifying core knowledge
- Explanations and worked examples
- Keywords highlighted in bold
- A set of quick check questions to check your understanding of the key concepts
- Answers to all questions

All of these photocopiable sections are designed to be used either in class, during a tutorial, during one-to-one sessions or by the students working alone in self-study. They will be particularly useful as a short preparation exercise before tackling exam-style questions. Although the primary focus is on securing a grade 5, Higher Tier students will find valuable support in these materials given the knowledge-rich nature of the examinations and looking to move from a solid grade 5 to a grade 6.

The activities are included within each chapter in non-write on format to reduce photocopying costs. However, for those teachers who prefer students to write on the activities, they are additionally available in this format as an appendix. Diagrams which are best printed in colour have been included as a PDF in the support files so that teachers can print a colour copy for their students.

January 2025

Chapter 1: Atomic structu
periodic table

1.1 Atoms, elements, compounds an

Background information

All substances are made of very tiny particles called **atoms**.

Substances that are made up of just **one type of atom** are called **elements**.

Elements are shown in the **periodic table**.

1 (1)	2 (2)													3 (13)	4 (14)
				Key			1.0 **H** hydrogen 1								
6.9 **Li** lithium 3	9.0 **Be** beryllium 4		relative atomic mass **atomic symbol** name atomic (proton) number											10.8 **B** boron 5	12.0 **C** carbon 6
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)			27.0 **Al** aluminium 13	28.1 **Si** silicon 14
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30			69.7 **Ga** gallium 31	72.6 **Ge** germanium 32
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	96.0 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48			114.8 **In** indium 49	118.7 **Sn** tin 50
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La**** lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium [illegible]	195.1 **Pt** [illegible]	197.0 **Au** gold 79	200.6 **Hg** mercury 80			204.4 **Tl** thallium 81	207.2 **Pb** lead 82
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac***** actinium 89	[267] **Rf** rutherfordium 104	[268] **Db** dubnium 105	[271] **Sg** seaborgium [illegible]	[272] **Bh** [illegible]	[illegible]	[illegible] meitnerium 109	[illegible] **Ds** darmstadtium 110	[280] **Rg** roentgenium 111	[285] **Cn** copernicium 112			[284] **Uut*** ununtrium* 113	[289] **Fl** flerovium 114

Each element is represented by a chemical symbol, e.g. **He** represents **one atom** o

Exam tip: You will be provided with a periodic table for your exam, you do not n

Note that the first letter of the symbol is always a capital letter and the second l

Most symbols use the first (and sometimes second) letter of the name of the ele
potassium (K) are exceptions.

Compounds contain **two** or more **different** types of **atoms**. Compounds are form
elements **react** and join to form **products**.

Worked example: word equation

hydrogen + oxygen ⟶

Compounds can be represented using **formulae**. Formulae are the **symbols** and n
in a compound

Hydrogen (H_2) and oxygen (O_2) exist as **molecules** (more than one atom). They ar
only one **type** of atom.

Water is a **compound** (two different types of atoms) and a **molecule** (more than o
water there are two hydrogen atoms and one oxygen atom, H_2O.

Worked example – chemical equation

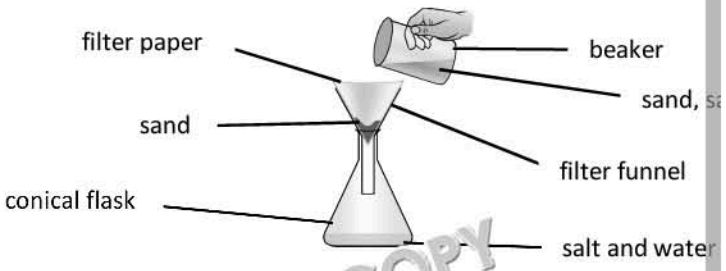
Using the formulae for hydrogen, oxygen and water, we can write a balanced eq

Exam tip: There are c... m...emicals that are used in many chapters throughc
familiar w... h ... names and the formulae. Note that the **numbers** are sub
and smalle... the letters.

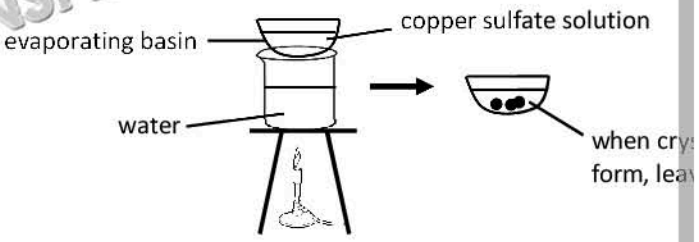
Mixtures contain two or more elements or compounds that are mixed. They have

Mixtures can be separated by physical processes:

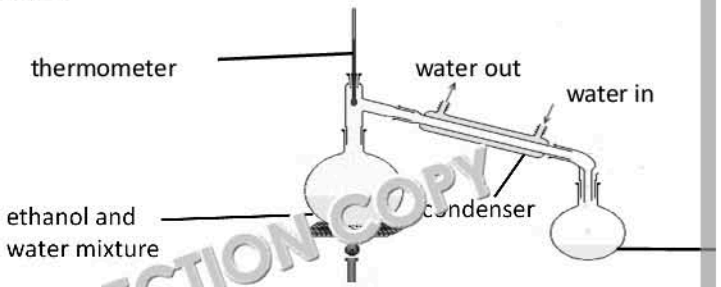
- **Filtration**; separates soluble solids/liquids from insoluble solids, e.g. salt wat



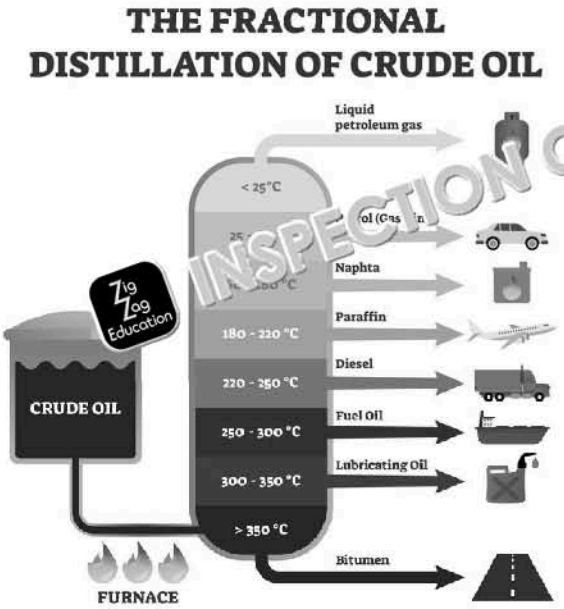
- **Crystallisation**; separates solu... id ...om a solution, e.g. copper sulfate s



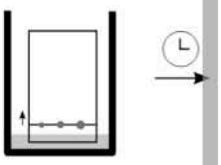
- **Simple distillation**; separates components in a mixture with a large differenc
e.g. ethanol and water



- **Fractional distillation**; separates components of a mixture with similar boiling points, e.g. crude oil



- **Chromatography**; components of a



Exam tip: Descr
Visualise/sketch
and use this ima
procedure. Bull
structure your a
Do not be afraid
answer the 6-m

Quick check 1.1

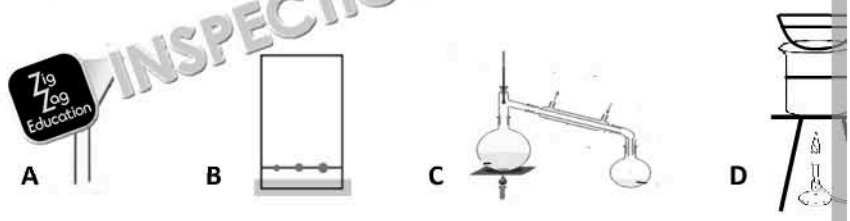
1. (a) Chlorine reacts with potassium bromide to produce bromine and potass
Write the word equation for this reaction
(b) Hydrochloric acid reacts with sodium hydroxide to produce sodium chlo
Write the word equation for this reaction.

2. Balance the chemical equations for the following reactions:
(a) ..... + ...Cl_2 → .....NaCl
(b) ..... + ...H_2 → .....NH_3
(c) .....Mg + ...O_2 → .....MgO

3. Use the Internet to research the formula for each compound given in the tab
Copy and complete the table. The first one has been done for you.

name	formula
water	H_2O
methane	
ammonia	
hydrochloric acid	
sulfuric acid	
carbon dioxide	

4. Which apparatus can be used to separate different food colourings? Select o



5. Describe how you would separate a mixture of salt and sand. Use diagrams t

1.2 The periodic table

Background information

Substances that are made up of just **one type of atom** are called **elements**.

A list of all known elements can be found in the periodic [illegible]

The development of the periodic [illegible] Mendeleev

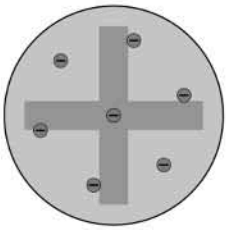
In early versions of the periodic [illegible] elements were placed in order of **atomic w** number). The problem [illegible] this were that the table was incomplete, and eleme inappropria[illegible]up

Mendeleev changed the order of the elements and left **gaps** for elements not yet

Discovery of **isotopes** – explained why atomic weight was not correct.

Models of the atom:

Thompson – plum pudding model. A ball of positive charge with negative electrons embedded in it.



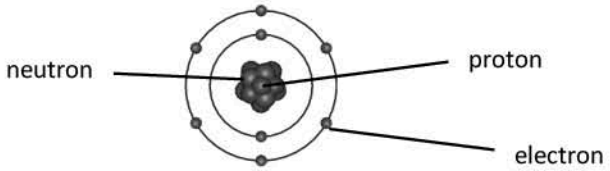
Rutherford – nuclear m The mass of an atom is concentrated at the ce (nucleus) and the nucle is charged.

Rutherford scattering experiment:

- positively charged alpha particles fired at thi[illegible]d
- most alpha particles went straight th[illegible] **conclusion** – atom mostly empty
- a few were scattered in diffe[illegible]ctions; **conclusion** – mass of an atom is

Chadwick – [illegible]nuclear model) evidence to show the existence of neutr

Bohr – (extension of nuclear model) electrons orbit the nucleus at specific distanc



Exam tip: Remember the names of these scientists and their theories/experimer time, and testing these theories can support or refute the ideas.

Nuclear model:

The currently accepted model of the atom is th[illegible] model; it consists of a nu neutrons, and electrons in orbits surr[illegible] nucleus.

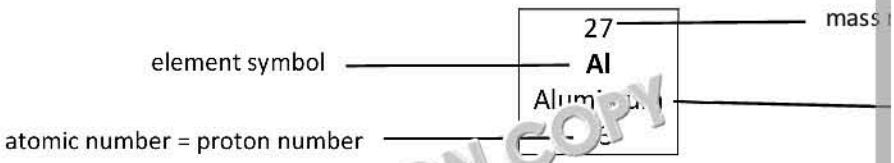
Subatomic partic[illegible]

[illegible]ticle	relative charge	
proton	+1	
neutron	0	
electron	-1	

Elements in the periodic table are currently arranged in order of **atomic (proton)**

Elements in the same group

- have the same number of **outer electrons**
- have similar **chemical** properties.



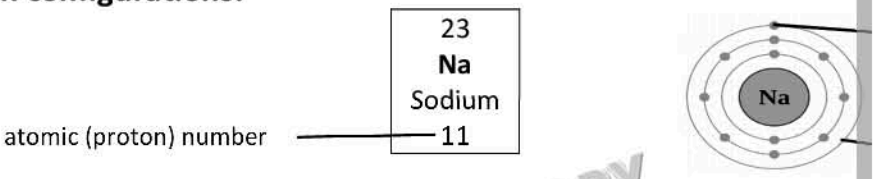
All **atoms** of aluminium (Al) have 13 **positive** protons. In an atom, number of elec
therefore a[tom] of aluminium also have 13 **negative** electrons and are **neutral**.

The atom of aluminium in the above example has 27 – 13 = 14 neutrons.

Exam tip: The higher number is always the mass number (protons + neutrons).

Isotopes: Atoms of the same element with the same number of protons but differ
Aluminium (Al) has two isotopes: $^{26}_{13}Al$ (13 protons and 13 neutrons) and $^{27}_{13}Al$ (1

Electron configurations:



Electron configurations can be represent[ed] in **diagram** or in **number** format. Th
diagram for sodium (Na) is shown above. The circles represent energy levels and
This can also be written in number format as 2,8,1.

Exam tip: Add 2 + 8 + 1 = 11 = number of protons (atomic number) = numb
of sodium.

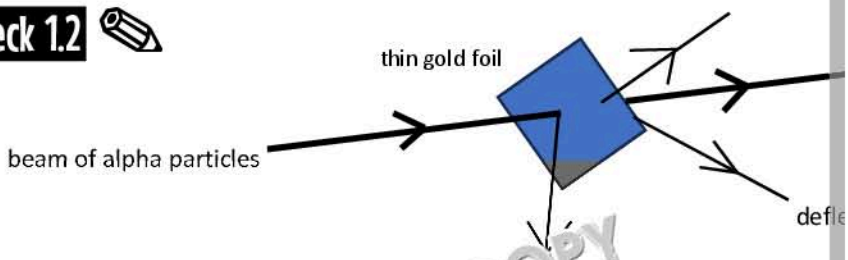
Drawing electron configurations:

- Write the element symbol in the centre of your paper and draw a circle arou
- Locate the element in the periodic table.
- Sodium is in **period 3**; therefore, draw three circles around your original circ
the last.
- Sodium is also in **group 1**. This means you should draw one dot somewhere
- Lastly, complete the inner circles. The **first circle** contains a maximum of **2** e
circles should contain **8** electrons (dots).

We can write the number format using the diagram we have drawn; just count th
circle working from the innermost circle to the outermost circle.
For sodium this is 2,8,1 (= 11 = atomic number).

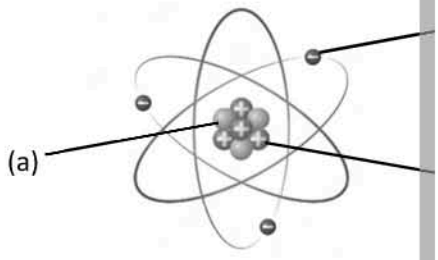
Exam tip: When drawing electron configuration diagrams, think of the **circles** as
the period table. The **first** circle represents **period 1**, which has **two** elements
maximum number of **electrons** (dots) that can be drawn is **2**. The **next circle** re
eight elements; therefore, the **maximum** number of **electrons** (dots) we can dra
us **how many** electrons to draw in the outermost circle, i.e. **group 6 = 6 electron**

Quick check 1.2



1. Alpha particles are fired at gold foil. Deflected alpha particles
   Write down two conclusions that can be made from these results.

2. Name the subatomic particles labelled (a) in the diagram to the right.



3. Copy and complete the element names and symbols that are missing from the periodic table to help you.

48 22	28 14	16 8	20 10

4. How many protons, electrons and neutrons do atoms of the following elements

24 **Mg** magnesium 12	4 **He** helium 2	**Ar** aluminium 13	

5. Figures 1 and 2 show Bohr model of an atom of two different elements in the periodic table.

   (a) What is the number of electrons and neutrons in **Figure 1**'s atom?
   What is the name of the element?

   (b) What is the number of electrons and neutrons in **Figure 2**'s atom?
   What is the name of the element?

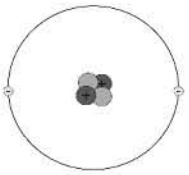


Figure 1

6. What is meant by isotopes? (You should refer to subatomic particles in your

7. Draw the electron configuration diagrams for an atom of each of the following

9 **Be** beryllium 4	31 **P** phosphorus 15	39 **K** potassium 19

8. Write down each electron configuration from Q7 in number format.

9. What did scientists use to arrange elements in early periodic tables?

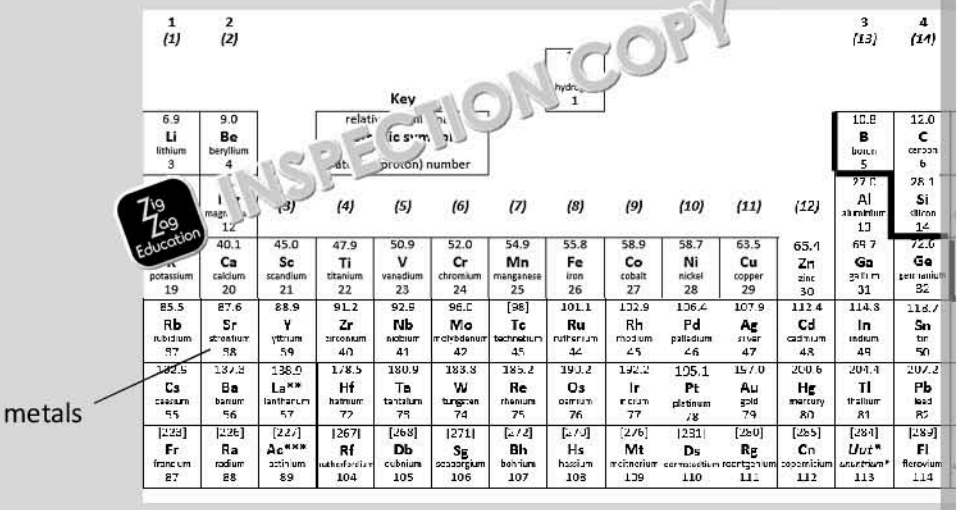
10. Complete the sentence.
    Mendeleev overcame the problem associated with the early periodic table by [redacted] had not yet been discovered.

11. State one similarity and one difference in the electronic structure of potassium

1.3 Groups, properties and trends

Background information

Metals and non-metals



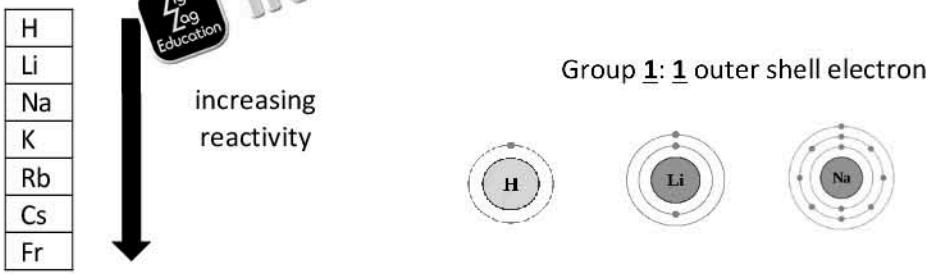
Metals react to form **positive ions**. Non-metals react with metals to form **negativ**

Group 1: alkali metals

Physical properties

- soft
- low melting points – melting points decrease down the group
- low densities

Chemical properties



Group 1 metals react by forming **positive** ions with a **+1** charge. They **lose** their si

Li **atom** Li **ion** (Li^+)
(full outer shell = sta

Reactions with oxygen

Alkali meta **to s under oil** as they **react** readily with **oxygen** in the air to fo
is seen and ction becomes **more vigourous** going **down** the **group**.

Worked example – word equation

metal	+	oxygen	→

Worked example – chemical equation

2M	+	$1/2O_2$	→	

Exam tip: Insert **any** group 1 metal name or symbol into the above equations. S
metal, e.g. M → Na (sodium).

Flame tests

Alkali metals also **burn** in o[illegible] to produce different coloured flames.

group 1 metal	fla
Li	
Na	
K	

Reactions with water

Alkali metals react vigorously with cold water to produce **metal hydroxide** and **hy**

Worked example – word equation

metal	+	water	→	metal hydro

Worked example – chemical equation

2M	+	$2H_2O$		2MOH

Exam tip: Substitute th[illegible] in the above equations for **any** group 1 meta

Observatio[illegible]

metal	add H_2O	add
Li	fizzes **steadily** (H_2 gas produced); **slowly** becomes smaller and eventually disappears	
Na	fizzes **rapidly** and forms a ball; **quickly** becomes smaller and disappears	
K	fizzes **violently**; sparks; disappears **rapidly** with small explosion	

Exam tip: You will be expected to describe the observations. The key describing
bold in the table.

Reactions with chlorine

Alkali metals react vigorously with chlori[illegible] roduce chlor**ide**s (white solids). Th
to form **colourless solutions**. The reactions become **more vigorous** going **down** t

Worked [illegible] word equation

metal	+	chlor**ine**	→	

Exam tip: Note the change in the ending of the halogen name from **ine** to **ide**.

Worked example – chemical equation

2M	+	Cl_2	→

Exam tip: Insert **any** group 1 metal in place of M into the above equations.

Explaining the trend in reactivity

Atoms **lose** their **1** outer electron to form **positively** charged ion. Going down g

- atoms become **larg**
- outer e… n … further from nucleus
- force o… **tion** between nucleus and outermost electron **decreases**
- electron is **easier** to lose.

Exam tip: Remember the observations; you can find clips of the reactions to wat

Group 7 – halogens (non-metals)

Physical properties

- exist as simple **molecules** Cl_2, Br_2, I_2
- **increase** ↑ in molecular **mass** going **down** the **group** ↓
- melting point and boiling point **increase** ↑ going **down** the **group** ↓

element	colour	
chlorine (Cl_2)	green	
bromine (Br_2)	br… n-r…d	
iodine(I_2)	purple	

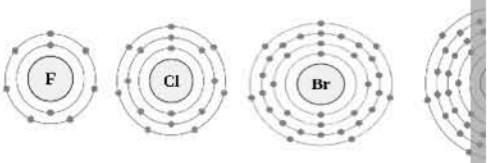
Exam tip: Note the **cha…** going down the group. Mass $I_2 > Br_2 > Cl_2$. H… **higher boil…** …nd … …g points as they have **stronger intermolecular** forces. I… overcome… …) these forces.

Chemical properties

F
Cl
Br
I
At

↑ Increasing reactivity

Group 7: 7 outer shell electrons



Group 7 halogens react by forming **negative** ions with a **-1** …rge. They **gain** a si…

Cl **atom** → Cl **ion** (Cl^-)

full outer shell = stabl…

Reaction with metals

The halogens react with metals to produce **ionic** salts.

Worked example – word equation

metal	+	chlor**ine**	→

Exam tip: You can replace the word 'chlorine' with any halogen and the word 'metal' with any metal from the periodic table and the word equation is the same.

Worked example – chemical equation

The chemical equations vary depending on the metal used. Reactions with group 1 metals are shown in the group 1 section. The equation for the reaction with a group 2 metal is shown below:

Me	+	Cl_2	→

Exam tip: You can replace **Me** with any **group 2 metal symbol** and the equation will be balanced. The chemical equation is different than for a group 1 metal. The reason why will be explained in the ionic bonding in Chapter 2.

Reaction with non-metals

When a halogen reacts with hydrogen a **hydrogen halide** is produced.

Worked example – word equation

hydrogen	+	brom**ine**	→

Exam tip: Again, the ending of the halogen name changes from **ine** to **ide**.

Worked example – chemical equation

H_2	+	Br_2	→

All hydrogen halides are **gases** at room temperature and they **dissolve** in water to form solutions.

Displacement reactions

A **more reactive** halogen can **displace** a **less reactive** halogen from **solutions of its salts**.

Worked example 1 – word equation

chlor**ine**	+	potassium brom**ide**	potassium chlor**ide**

Worked example 2 – chemical equation

Br_2	+	2KI	→	2KBr

Exam tip: Remember reactivity **decreases down** the **group** (Cl > Br > I). Chlor**ine** will displace an iod**ide**. Brom**ine** will displace an iod**ide**. The name endings change.

Explaining the trend in reactivity

Atoms **gain** 1 outer electron to form a **negatively** charged ion. Going down group

- atoms become **larger**
- outer electron is **further** from nucleus
- force of **attraction** between nucleus and outermost electron **decreases**
- electron is **harder** to gain.

Exam tip: The explanations for trend in reactivity of group 1 and group 7 are ver electron and group 7 **gain** 1 electron. Atoms lose electrons more easily a electrons for the exact same reason!

Group 0 - the noble gases (non-metal)

Physical properties

- exist as single **atoms**
- low boiling points – boiling points increase ↑ going down group ↓

Chemical properties

He
Ne
Ar
K
Xe

inert (unreactive)
full outer shell (2 (He) or 8 electrons)
stable electron configuration

Quick check 1.3

1. Lithium, sodium and potassium are group 1 elements. What is the name give
2. Lithium reacts with water to produce lithium hydroxide and hydrogen. Copy equation for this reaction:

   ..... Li + ... H_2O → ..... LiOH + ...

3. Group 1 metals all react with water. Compare what is seen when lithium rea reacts with water.
4. Sodium reacts with chlorine to produce sodium chloride.
   (a) Write a word equation for this reaction.
   (b) Describe what happens when a lithium atom reacts with a chlorine atom transfer in your answer.
5. An atom of chlorine is represented as $^{35}_{17}Cl$
   (a) Determine the number of protons and the number of neutrons in one at

   As well as $^{35}_{17}Cl$, chlorine atoms can be of the form $^{37}_{17}Cl$

   (b) What is the name given to these different atoms of chlorine?
6. Which **types** of element react when Mg reacts with Cl?
7. The halogens Cl_2, Br_2 and I_2 all react with silver. Which halogen is the **least** re Explain your answer.
8. Give two trends in the physical properties of the halogens.
9. Chlorine reacts with potassium iodide solution.
   (a) Complete the word equation below.

   chlorine + potassium iodide → ..... + ...

   (b) What is the name of this type of reaction?
10. (a) What is the name of the group that contains neon?
    (b) What is the electronic structure of a neon atom?
    (c) Explain why **no** products are formed when neon is burned in oxygen.

Checklist: Atomic structure and the periodic ta

I can...
understand the difference between an atom, an element, a compound and a mi
use the names and symbols of the first 20 elements in the periodic table (H to Ca
know the names and formulae of some simple chemical compounds
write word equations for simple chemical reactions
balance symbol equations for some simple chemical reactions
describe how to separate simple mixtures
understand how scientific models and theories develop over time
describe the scattering experiment and draw conclusions from the results
describe the difference between the plum pudding model and the nuclear mode
locate elements in the periodic table
calculate the number of protons, neutrons and electrons in an atom
describe the atomic structure of an atom in terms of its protons, neutrons and e
represent the electronic structure of elements in the periodic table in both numb and diagrams
describe the steps in the development of the periodic table
balance simple chemical equations for the reactions of group 1 (alkali metals) an elements
explain how the physical properties of group 1 (alkali metals), group 7 (halogens gases) elements depend on the outer shell electrons
explain how the chemical properties of group 1 (alkali metals), group 7 (halogens gases) elements depend on the outer shell electrons
describe what will be observed when the first three group 1 (alkali metals) elem oxygen, chlorine and water
describe the nature of the compounds formed when group 7 (halogens) element and non-metals
predict properties of group 1 (alkali metals), group 7 (halogens) and group 0 (no trends down the group

Chapter 2: Structure and

There are **three** types of **strong** chemical bonds:

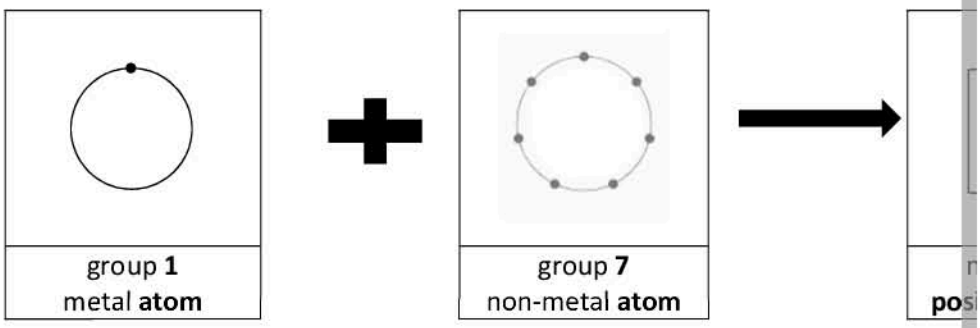
- **ionic**
- **covalent**
- **metallic**

2.1 Ionic bondi

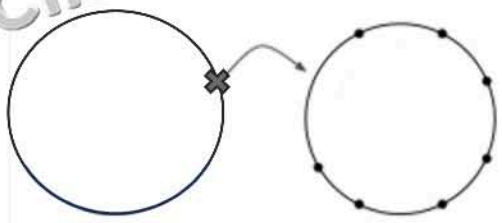
Backgrou fo

Formed bet metals and **non-metals.**

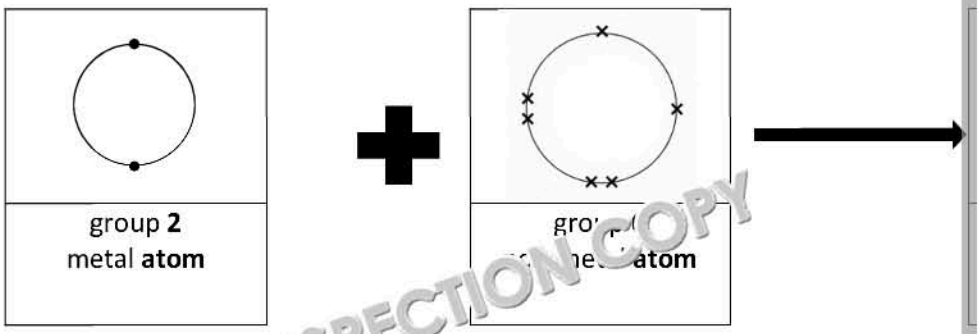
Worked example



Note the **transfer** of 1 electron from the outer shell of a gr 1 **atom** to the oute
Both **ions** now have a **stable electron configuratio** roup 1 **ion** will have **0** e
(sometimes shown as **2** or **8** depending on t r 1 metal). The group 7 **ion** w
outer shell.



Worked example



Note the **tra** o ons from the outer shell of a group 2 **atom** to the out
ions now h **able electron configuration**. The group 2 **ion** will have **0** electr
shown as **2** depending on the group 2 metal). The group 6 **ion** will have **8** ele

Exam tip: The **metal** atom will always **lose** electrons and form a **positive** charge
always **gain** these electrons and form a **negative** charged **ion**. The **number** of e
equivalent to the **charge** on the **ion**.

Ionic compounds

- **giant** structure of **oppositely charged ions**
- **ions** held by **strong** electrostatic forces of **attraction** which act in **all directio**
- **high** melting and boiling points – **large** amounts of **energy** required to **break**
- **do not** conduct electricity when **solid** because **ions** are **fixed**
- conduct electricity when **molten** or **aqueous** because **ions** are **mobile**

Ionic models

space-filling

ball and stick

Representing ionic structure using models has advantages but it also has limitatio

model	advantage	li
space-filling	clearly shows the arrangement of **ions** in **one** layer	no information abo does not sh
ball and stick	shows arrangement of **ions** in a larger section of the crystal than space-filling model	no information abo sticks for b hows free space be
dot and cross	shows **transfer** of electrons	does not show ho

Exam tip: You will be exp[illegible] recognise **models** of ionic compounds. The **po**
The **negative ion** i[illegible] the non-metal. To work out the **empirical formula** co
e.g. Na_5Cl_5 [illegible] is equivalent to Na_1Cl_1 (NaCl).

Quick check 2.1

1. The diagram shows a dot and cross diagram for potassium fluoride.

What type of bonding is present in potassium fluoride?

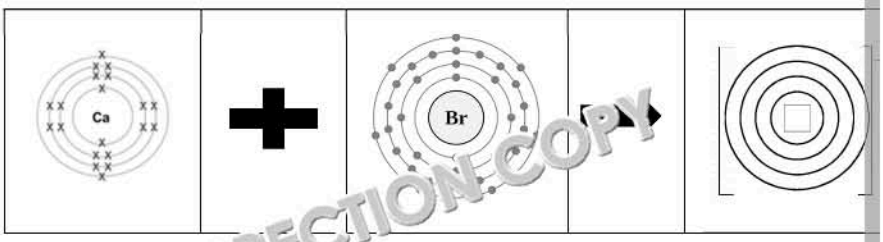
2. Complete gaps **(a) – (f)** in the paragraph below. Use words from the box.

strong	**melting**	**molten**	**attraction**	**mobile**

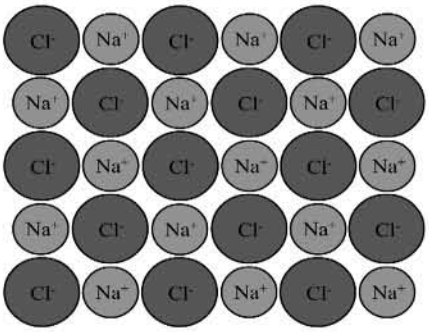
*Beryllium oxide forms a **(a)** structure of oppositely charged ions (ionic lattice) **(c)**. It has a high **(d)** point and conducts electricity when in **(e)** state due to **(f)***

3. Describe how a calcium atom and a sulfur atom would react to form calcium sulfide. Include the structure of each atom and the electronic structure and charges of any ions formed.

4. Calcium reacts with bromine to produce calcium bromide. Copy and complete the diagram from this reaction.



5. The diagram shows the structure of an ionic lattice. Which force holds the particles together?

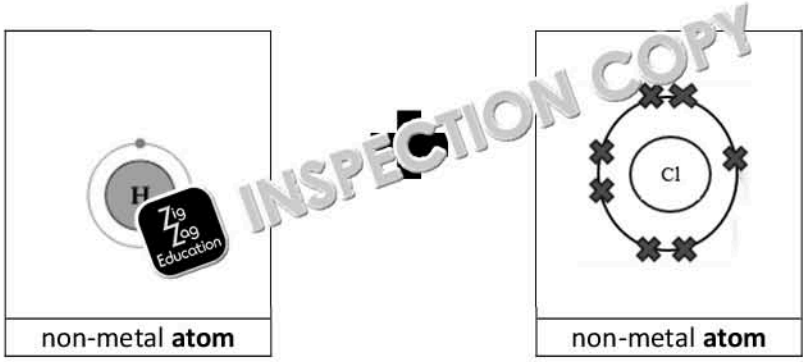


6. Solid ionic compounds do not conduct electricity.
   (a) Explain why.
   (b) How can you make an ionic compound conduct electricity?
   (c) Explain how this makes the ionic substance conduct electricity.

2.2 Covalent bonding

Generally formed between **non-metals**.

Worked example

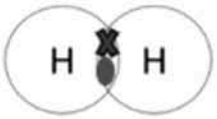
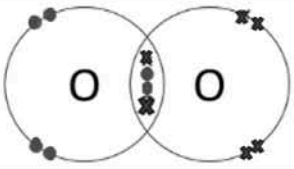
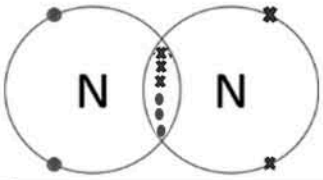


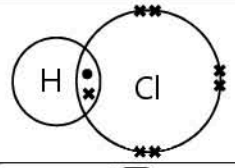
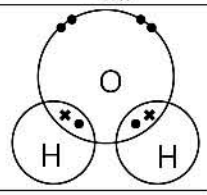
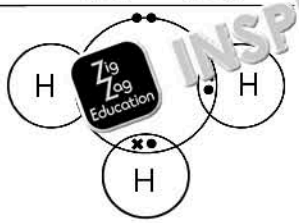
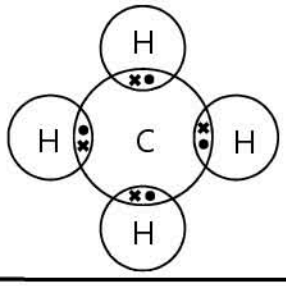
non-metal **atom** | non-metal **atom**

Note the **sharing** of **2 electrons** to form the **covalent** bond. This **bond** is represen
outer electron shells. **One** electron is from the outer shell of **hydrogen** (dot) and
shell of **chlorine** (cross). Both **atoms** now have a **stable electron configuration** a

Background information

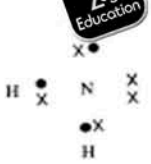
Dot and cross diagrams for small covalent molecules

	**Hydrogen** 2 hydrogen **atoms** react to form a hydrogen **molecule**, H_2
	**Chlorine** 2 chlorine **atoms** react to form a chlorine **molecule**, Cl_2
	**Oxygen** 2 oxygen **atoms** react to form an oxygen **molecule**, O_2
	**Nitrogen** 2 nitrogen **atoms** react to form a nitrogen **molecule**, N_2

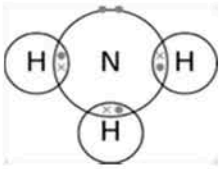
	**Hydrogen chloride** **1** hydrogen **atom** reacts with **1** chlorine **atom** to form a hydrogen chloride **molecule**, HCl	**Exam tip** These are and cross **different** bond to a Cl is in gr complete forms a s O is in gr complete forms a s N is in gr complete forms a s C is in gr complete forms a s
	**Water** **2** hydrogen **atoms** react with **1** oxygen **atom** to form a water **molecule**, H_2O	
	**Ammonia** **3** hydrogen **atoms** react with **1** nitrogen **atom** to form an ammonia **molecule**, NH_3	
	**Methane** **4** hydrogen **atoms** react with **1** carbon atom to form a methane **molecule**, CH_4	

Exam tip: You will be expected to complete dot and cross diagrams for the molecul
expected to draw **outer** shell electrons. If the atom is bonded to **H** atom(s) you wil
1 dot and **1 cross** in the crossover of the circles. This for **every** H atom and **do n**
the H atom. You can now draw electrons on the **other atom** to make the number

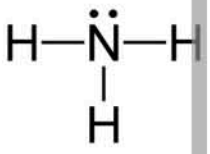
Covalent models



dot and cross
(without circles)



dot and cross
(with circles)



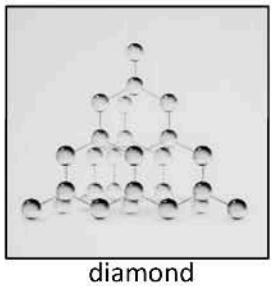
displayed formula
(2D)

As for ionic structures, representing covalent structures using models has advanta

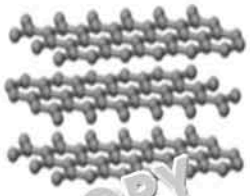
model	advantage	
dot and cross	shows **sharing** of electrons shows **lone pairs** of electrons	does not show
2D displayed formula	shows that the atoms are connected via **bonds** shows **double**, **triple** bonds	does not show
3D ball and stick	shows how the **atoms** are connected shows how **atoms** are arranged in space shows the angles between bonds	no information sticks f shows **gaps** bet

Exam tip: You will be expected to recognise models of covalent compounds. To
formula count the number of each **atom**, e.g. $N_1H_3 = NH_3$

Giant covalent structures: examples



diamond



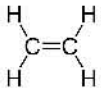
graphite

Silica and d[illegible] have the same structure.

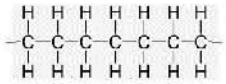
- In **diam[illegible]** all the atoms are **carbon** and there is an extensive network of bo[illegible] Each carbon atom is bonded to **4** other carbon atoms.
- In **silica** the atoms are **silicon** and **oxygen**. **Oxygen** is the **smaller** atom. Each[illegible] oxygen atoms in an extensive network. Each oxygen atom is bonded to **2** sili[illegible]

Polymers: examples

Polymers are **long-chain** covalent molecules.



ethene
monomer
(small molecule)



polyethene
polymer
(long-chain m[illegible]lecule)

repe[illegible]
n

Many **small** ethene monomers will **join** to fo[illegible] ti[illegible] polymer. The double bond br[illegible] with a second monomer. This hap[illegible] [illegible]ously and a **long chain** is formed.

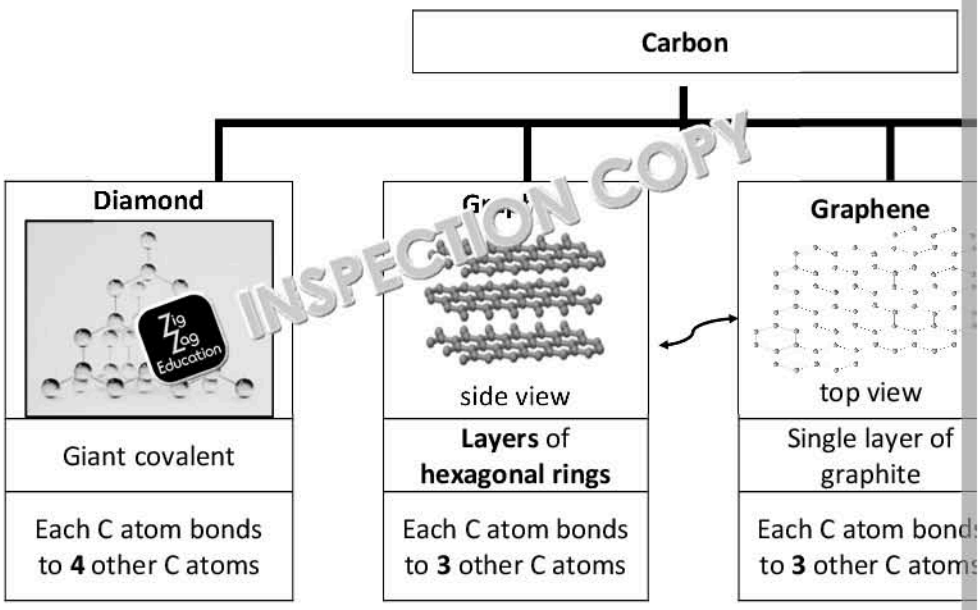
Exam tip: Yo[illegible] to draw polymers, just recognise the structures.

Covalent co[illegible]ounds

- **share** electrons
- can be **small** or **giant** or **long-chain molecules (polymers)**

molecule size	small	giant
state	gas or liquid	solid
melting and boiling point	low	high
forces/bond broken	weak intermolecular forces	strong covalent bonds
energy needed to break bond/force	low	high
conduction of electricity	do not cond[illegible]ct zero m[illegible]ge[illegible]a[illegible]ged [illegible]a[illegible]ticles	

Structure and bonding of carbon
Carbon has **four** different **covalent** structures.



Fullerenes – based on hexagonal rings.


C_{20}


C_{60}

Exam tip: You will be asked to recognise the structure and bonding in these g
them to th op

carbon structure	physical property		
	general	melting point	electrical c
diamond	**hard** – strong covalent bonds	**high** – giant structure **strong** covalent **bond broken**	no – zero charged p
graphite	**soft/slippery** – **layers slide** over each other	**low – weak** intermolecular **forces between layers**	yes – **delo mobile** el
graphene	**strong** – strong covalent bonds	**High** – **strong** covalent **bonds broken**	yes – **delo mobile** el
fullerene	**hollow shape**, e.g. tubes, balls, cages		yes – nanotu **delocalised** electr

Exam tip: Note the difference in the melting point of graphite. **Melting** points a
atoms) ar en, and are **low** if intermolecular **forces** (between **molecules/la**

Quick check 2.2

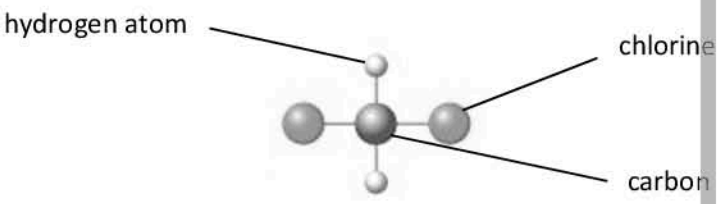
1. The diagram on the right shows a model of a nitrogen molecule. What do the lines between the atoms represent?

2. (a) Copy and complete the dot and cross diagram on the right for a molecule of water. Show only the electrons in the outer she

(b) The diagram bel of a water molec Give **one** limitat model of water

(c) Give **one** limitation of using the dot and cross model of water over the b

3. Give the molecular formula of the molecule below.



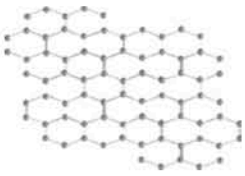
4. Sulfur is below oxygen in the periodic table.
   (a) What type of bonding is present in molecules of sulfur?
   (b) Deduce the dot and cross diagram for a molecule of sulfur, S_2. Show only
   (c) Complete the sentences by choosing the correct option of the words in
   The boiling point of sulfur is **(a)** ***low/high*** *because sulfur contains* ***between its molecules.***

5. The diagram below sh acture of C_{60}.

Select the correct answer. C_{60} is...

giant ionic	**graphite**	**metallic**

6. Graphene is a single layer of graphite.



What type of bonding is present between the carbon atoms in graphene?

7. Complete the sentences.
   The atoms in a polymer molecule are held tog er by **(a)** *bonds. The polyme forces. When a polymer melts or boils, th* **(c)** *are broken.*

8. Complete the sentenc
   (a) *In graphit arbon atom forms bonds with* **(i)** *other carbon atoms. d* **(i)**.
   (b) Gi **other** physical properties of graphite.
   (c) Suggest why graphite can be used as a lubricant.

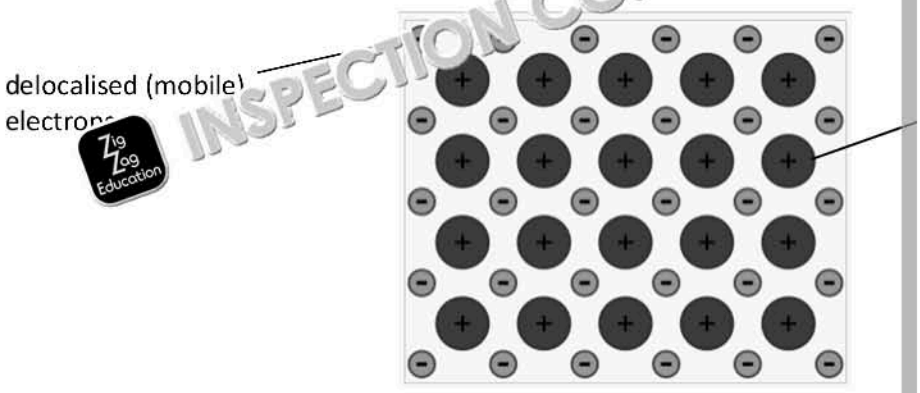
9. Complete the sentences.
   The structure of a fullerene is based on **(a)**. *The fullerene molecule is made fr One use of the fullerene molecule is* **(c)**.

2.3 Metallic bonding

Background information

Giant structures of **metal atoms** arranged in a **regular pattern**.

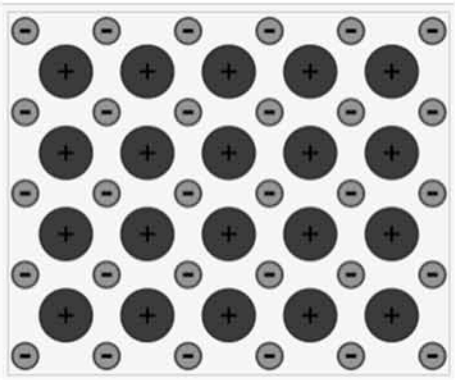
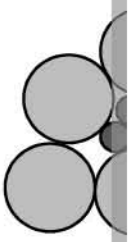
Electrons in outer shell of each atom **delocalise** and **mov** oughout structure.



properties	
good electrical conductors	**delocalised electrons** carr
good thermal conductors	**delocalised electrons** trans
high melting and boiling points	**strong** met **high** e
bent and shaped easily	

Alloys

Mixtures of two or ma ts – at least **one** is a **metal**.
Metals are d t make them **harder**.

pure metal	
	
atoms arranged in **layers**	**atoms**
layers can **slide**	**layers** disto
	harde

Quick check 2.3

1. Explain how alloying a metal improves its properties. Draw diagrams to help
2. Explain why a copper alloy is harder than pure copper metal.

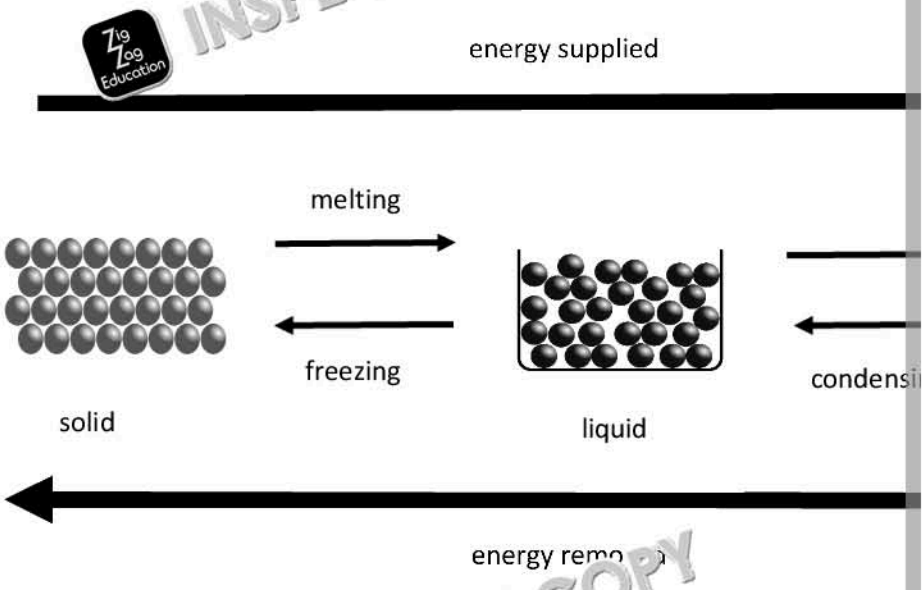
2.4 Bonding, structure and properti

Background information

The **three** states of matter are:

- **solid**
- **liquid**
- **gas**

These **three** states of matter c[illegible] nte[illegible]ange depending on the amount of energy



The amount of energy required for a substance to change state depends on the st
the particles.

Exam tip: [illegible] and boiling points

type of compound	force is between	what is broken	stron
ionic	ions	electrostatic attraction (ionic bond)	str
small covalent	molecules	intermolecular forces	we
giant covalent	atoms	covalent bond	str
metals	atoms	metallic bond	str

The odd one out is the **small covalent molecules**. To melt or boil these molecul
between molecules. These are **weak** and require **low** amounts of **energy**. To m
bonds between atoms/ions are broken. These are **stro**[illegible]d require **high** amo

State symbols

In chemical equations t[illegible]e states of matter are shown as (s), (l) and (g). You
means that [illegible]bs[illegible]ce is dissolved in water and is therefore an **aqueous** soluti

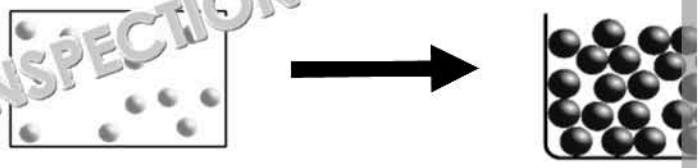
Exam tip: Note the **brackets** and **small letters** used as opposed to capitals. We
chapters. Water is always $H_2O(l)$.

Quick check 2.4

1. Copy and complete the chemical equation using state symbols. Sodium meta solution of sodium hydroxide and hydrogen gas.

$$Na(....) \quad + \quad H_2O(....) \longrightarrow NaOH(....)$$

2. A gaseous molecule is cooled to form a liqui[d] d[u]ring an experiment. Name t[h]



3. The melting point of lithium is 180 °C. What is the state of lithium at 150 °C?

4. (a) What is the general term used for the letter(s) given in brackets in an eq or compound?
   (b) What does (aq) mean?

Checklist: Structure and bonding

I can...
name the three types of strong chemical bonds
explain ionic bonding in terms of transfer of [e]lec[tr]o[n]s, formation of ions and str[u]
complete dot and cross diagra[m]s to [r]epresent ionic compounds and some small
deduce that a comp[ou]nd is ionic or covalent from a diagram of its structure
describe th[e] limitations of using dot and cross, ball and stick and space-filling mo ionic compounds and covalent compounds
work out the empirical formula of an ionic compound from a model or diagram
explain covalent bonding in terms of sharing of electrons and formation of strong
work out the molecular formula of a covalent compound from a model or diagra
explain metallic bonding in terms of giant structure of atoms, arranged in regular delocalised electrons moving throughout structure
recognise metals and alloys from diagrams
compare physical properties of metals and alloys in terms of structure and bondi
state and explain the physical properties of ionic and [co]valent (small and giant) c in terms of their structure and bondi[ng]
explain changes of state in terms of energy transfer and types of bonding
recognise [and] include appropriate state symbols in chemical equations
recognise different structures of carbon from diagrams or descriptions
explain the properties of the different structures of carbon in terms of their stru
give examples of the uses of different structures of carbon based on their proper

Chapter 3: Quantitative c

3.1 Balanced equations and conser

Background information

Chemical equations provide a way of represent[illegible] chemical reactions and commu

Word equation

[illegible]ydrogen	+	oxygen	→	

Symbol equation

$2H_2$	+	O_2	→	

Symbol equations must be **balanced** in terms of the number of atoms of each ele the equation. This is due to the law of conservation of mass.

Left

atom (reactants)	number
H	4
O	2

Right

atom (product
H
O

The law of conservation of mass states that no atoms are lost or made during a ch

[illegible] products = mass of reactants

This rule can appear [illegible] true when **gases** are produced as the gas may escape

Worked example

24 g of magnesium reacted with 71 g of chlorine to produce magnesium chloride. chloride was produced?

Mass produced = 71 + 24 = **95 g**

Quick check 3.1

1. Potassium reacts with water to produce potassium hydroxide and hydrogen. Copy and balance the equation for this reaction.

.....K	+	[illegible]O	→	2KOH

2. 4.1 g of potassium [illegible] with fluorine to produce 9.6 g of potassium fluorid that re[illegible]

3. Solid calcium carbonate ($CaCO_3$) decomposes to produce calcium oxide solid (CO_2). 100 g of calcium carbonate was heated during an experiment and 56 g
   (a) Why is the mass of calcium oxide **less** than the mass of calcium carbonat
   (b) What mass of carbon dioxide is formed?

3.2 Relative atomic mass (A_r)

Background information

A_r = the average mass of atoms of an element that takes into account th

This can be found using the periodic table (the number above the element).

Worked examples

12 **C** carbon 6	16 **O** oxygen 8

Relative atomic mass (A_r): 12 16

We can calculate **relative atomic mass** of an element from the percentage abund

Worked example

The mass numbers and percentage abundance of two isotopes of Cl are shown.

mass number	percentage abund
35	75
37	25

Calculate the relative atomic mass of chlorine.

$$\frac{(35 \times 75) + (37 \times 25)}{100} = \mathbf{\underline{35.5\ g}}$$

As you can see from the worked example, mass number will always be a whole nu number – it depends on the percentage abundance of each isotope.

Quick check 3.2

1. The mass numbers and percentage abundance of three isotopes of oxygen a

mass number	percentage abund
15	15
16	50
17	35

Calculate the relative atomic mass of oxygen.

3.3 Relative formula mass (M_r)

Background information

M_r = the sum of the relative atomic masses of the atoms in the numbe

Worked example

Calculate the relative formula mass of $C_{\ldots}$

A_r: Cu = 63.5, C = 12, O = 16

$M_r = 63.5 + 12 + (3 \times 16) = \underline{\mathbf{123.5\ g}}$

The percentage by mass of an element in a compound is found using the followin

$$\frac{\text{relative atomic mass} \times \text{number of atoms}}{\text{relative formula mass}} \times 10$$

Worked example

Magnesium reacts with chlorine to produce magnesium chloride, $MgCl_2$. Calculat magnesium in magnesium chloride.

A_r of Mg = 24, M_r of $MgCl_2$ = 95

$$\frac{24 \times 1}{95} \times 100 = \underline{\mathbf{25.3\ \%}}$$

Quick check 3.3

1. Calculate the relative formula mass of $MgSO_4$.
   A_r: Mg = 24, S = 32, O = 16

2. The relative formula mass of $\mathbf{X}CO_3$ is 123.5
   Relative atomic masses A_r: C = 12, O = 16
   (a) Calculate the relative atomic mass of **X**
   (b) Name X

3. Magnesium reacts with oxygen to produce magnesi... oxide, MgO. Calculat oxygen in magnesium oxide (A_r: O = 16, M_r: MgO = ...).

3.4 Chemical measurements

Background information

Whenever a measurement is made, there is always some **uncertainty** in the result

The uncertainty in a set of measurements is ± half the range

A thermometer with a mark every 1.0 °C has an uncertainty of ± 0.5 °C.

A measuring cylinder with a mark every 1.0 cm^3 has an uncertainty of ± 0.5 cm^3.

The **mean** of a set of measurements is the average of all the numbers.

mean = sum of all measurements / number of measurements

The **range** of a set of measurements is defined as

range = biggest measurement – smallest measurem

Worked example

Calculate the mean volume and the uncertainty for the experiment below.

experiment	1	2	3	
volume (cm^3)	25.5	24.0	24.0	

$$\text{Mean} = \frac{25.5 + 24.0 + 24.0 + 26.0 + 23.0}{5} = 24.5 \text{ cm}$$

Range = 26.0 – 23.0 = 3.0 cm^3

$$\text{Uncertainty} = \frac{3.0}{2} = 1.5 \text{ cm}^3$$

Final answer = 24.5 cm^3 ± 1.5 cm^3

Quick check 3.4

1. **Table 1** shows the results of three experiments when solid calcium carbonate calcium oxide solid and carbon dioxide gas.

	mass of calcium oxide produced (g)	
1	2	3
54.3	55.4	56.1

Table 1

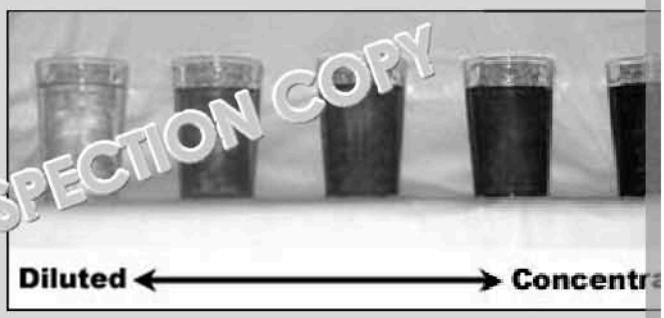
(a) What is the mean value?

(b) Calculate the uncertainty in the range of measurements.

3.5 Concentration of solutions

Background information

The **concentration** of a solution can be measured in mass per given volume, e.g. g



The more concentrated the solution, the more grams of solute that are dissolved

Worked example

A copper sulfate solution consists of 500 g/dm^3 of $CuSO_4$ dissolved in 1 dm^3 of w

Calculate the mass of solid copper sulfate used in 50 cm^3 of the solution.

(NOTE 1 dm^3 = 1000 cm^3) 500 g in 1000 cm^3

We have 50 cm^3 of solution, therefore we must divide the original 1000 cm^3 by 5

$$\frac{1000}{50} = 20$$

50 cm^3 is 1/20 of 1000 cm^3, therefore we need t[illegible]e 1/20 of the original 5

$$\frac{500}{20} = \mathbf{25\ g}$$

Exam tip: e[illegible]m you may be expected to rearrange an equation to calcul[illegible] familiar wi[illegible]w to rearrange an equation. You may also be given information the exam, e.g. gold (Au). The numbers may be different but the calculations foll[illegible] are allowed a calculator.

Quick check 3.5

1. A copper chloride solution contains 6 g of copper chloride in 200 cm^3 of solut[illegible] copper chloride in 50 cm^3 of this solution.

Checklist: Quantitative chemistry

I can...
balance chemical symbol equations
calculate the reacting masses in [illegible]ment
calculate the relative [illegible]ss of an element when given the percentage abu[illegible]
calculate t[illegible]tive formula mass of a compound
calculate the percentage by mass of an element in a compound when given the [illegible]
calculate the mean and uncertainty when given the values in an experiment
explain the changes in mass in an experiment
calculate the mass of a solute in a given volume of solution of known concentrat[illegible]

Chapter 4: Chemical ch

4.1 Reactivity of metals

Background information

Oxidation

Metals react with oxygen to produce metal oxides. In this reaction the metal **gain**

Word equation

Reduction – extraction of metals using carbon

Most metals are found in Earth's crust combined with other elements such as oxy contains enough of the metal compound to make it worthwhile extracting.

If a metal is **less reactive** than **carbon** (see below), it can be extracted from its ore

Word equation

In this reaction, carbon is **oxidised** as it **gains oxygen** from the metal oxide. The m **loses oxygen**.

The reactivity series

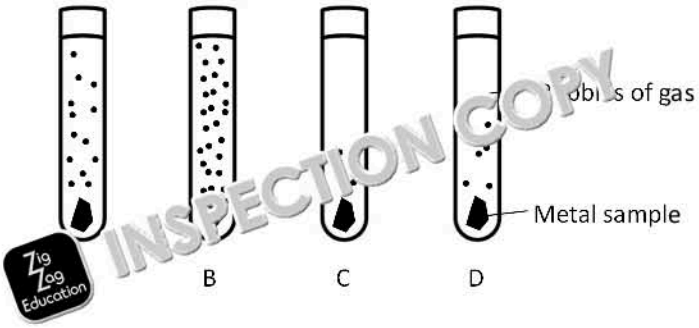
Metals can be arranged in order of their reactivity. The reactivity of a metal is rela positive ions.

	reaction with water (at room temperature)	reaction with dilute ac
potassium	very violent	very violent
sodium	violent	very violent
lithium	less violent	violent
calcium	fast	rapid
magnesium	slow	rapid
carbon		
zinc	no reaction	slow
iron	no reaction	slow
hydrogen		
copper	no reaction	no reaction

Carbon and hydrogen are not metals but are included in the reactivity series. Me reactivity series **do not react with water** at room temperature. Metals **below hyd not react with dilute acids**.

Worked example

Four metals were placed in separate test tubes containing a dilute acid.



Place the metals in order of their reactivity from most reactive to least reactive.

Answer: B, A, D, C

Exam tip: You will be expected to recall and describe the reactions of the above dilute acid. When a reaction takes place, a gas (hydrogen) is given off; therefore You will also be expected to place these metals in order of reactivity. Use a mne reactivity of the metals, e.g. **p**eople **s**ay **l**ittle **c**hildren **m**ust **c**hange **z**ebras **i**nto h

Reaction of metal with water

Word equation

metal + **wa** → **metal hydroxide**

In this reaction m hydroxide is formed. Metal hydroxides are **alkaline** (see b

Reaction of metal with dilute acid

Word equation

metal + **acid** → **metal salt**

This reaction produces a **metal salt** which is **neutral**. A metal **chloride** is produced **Sulfuric** acid would produce a metal **sulfate** and **nitric** acid would produce a meta

In both reactions (water and dilute acid), hydrogen gas is produced and therefore The **more reactive** the metal, the **more violent** the bubbling.

Worked example

Calcium sulfate can be produced by reaction between acid with a metal. Name produce calcium sulfate.

Acid: Sulfuric acid – this comes from the sulfate.
Metal: Calcium – the metal is always the first word of the metal salt.

Displacement reactions

We can deduce the order of reactivity of metals based on the results of displacem
metal can **displace** a **less reactive** metal from its compounds.

Worked example

Three metals, X, Y and Z, were added to separate solu[illegible] zinc chloride, and th
change measured.

metal	X	Y
temperatur[illegible]	9	no change

Use the reac[illegible] series to suggest identities for metal X, Y and Z. Give a reason f

Answer: Metals Z and X both displace zinc from the zinc chloride solution and the
reactivity series. Metal Z has the biggest temperature change and therefore must
the reactivity series. Metal Z could be lithium and Metal X could be magnesium.

Metal Y does not displace zinc from the zinc sulfate solution and therefore must b
series. Metal Y: iron or copper.

Exam tip: The key to answering questions about displacement reactions is to kn
the reactivity series. Metals **above** the metal contained in the solution will **displ**
will take place causing a **temperature change**. The **more reactive** the metal, the
change. Metals **below** the metal contained in the solution will **not displace** it, a
occur. Highlight the metal contained in the solution and then use your mnemon
series in the correct order.

Quick check 4.1

1. (a) A[illegible] reacts with zinc to produce zinc nitrate and hydrogen. Which ac
zinc nitrate?
   (b) What type of substance is zinc nitrate?

2. Calcium and an unknown metal X were added to separate solutions of zinc su
was recorded.

metal	temperatu
calcium	
X	

Suggest the name of metal X.

4.2 Reactions of acids

Background information

Neutralisation

Acids are **neutralised** by **alkalis/bases** to produce a **salt** and **water**.

Word equation

acid + alkali → salt

Acids are neutralised by **metal carbonates** to produce **salt, water** and **carbon dioxide**.

Word equation

acid + metal carbonate → salt + [illegible]

The pH scale

The pH scale is a measure of the acidity or alkalinity of a solution.

0	1	2	3	4	5	6	7	8	9	10
red			orange → yellow				green	light blue		
acidic										
most ←						least	neutral	least →		

The pH scale can be measured using **universal indicator** and a colour chart or a pH

Acids produce **H^+ ions** in aqueous solutions (pH 0–6) and **alkalis** produce **OH^- ions**

Universal indicator turns orange/red in acidic solutions, blue/purple in alkaline so
neutral solu

Symbol equation (acid)

$HCl(aq) \rightarrow H^+(aq) + Cl^-$

Symbol equation (alkali/base)

$NaOH(aq) \rightarrow Na^+(aq) + OH^-$

In neutralisation reactions, the H^+ ions (from the acid) react with the OH^- ions (fro
water (H_2O). The pH of the solution will be approximately 7 (neutral).

Symbol equation

$H^+(aq) + OH^-(aq) \rightarrow$

Worked examples

1. The pH of a solution is 9. Some sulfuric acid is added to the solution. Sugges

   Answer: pH = 6 (any number below 9 is acceptable as sulfuric acid would ma
   hence lowering the pH)

2. Describe a **method** that could be used to identify which of the solids, A, B or
   (NaOH), which is sulfuric acid (H_2SO_4) and which is so[illegible]n sulfate (Na_2SO_4).

solid	pH when
A	

Answe

1. Dissolve solid A in water.
2. Add universal indicator into the solution.
3. Record the result.
4. Repeat for solids B and C.
5. i) Solution A will turn purple – this is sodium hydroxide as it is an alka
   ii) Solution B will turn red – this is sulfuric acid as it is an acid.
   iii) Solution C will turn green – this is sodium sulfate as it is a neutral sa

Soluble salts

Soluble salts can be made by reacting acids with **metals** or **metal oxides/hydroxi**

Practical activity – Preparation of a pure, dry, soluble salt from an insoluble soli

1. **Dissolve** powdered **solid** in **acid** in a beaker.
2. **Stir** the reaction mixture and continue **adding** solid **until** it is in **excess**.
3. **Filter** the mixture to remove excess solid.
4. **Heat gently** in an evaporating dish until crystals [illegible]n to appear (saturated).
5. **Leave** the saturated solution in a [illegible]ce for a few days.
6. **Dry** the crystals between [illegible] sheets of filter paper.

Exam tip: [illegible] will [illegible]pected to recall this method of preparation so make sure
correct or[illegible]**chloride** crystals are made, e.g. copper chloride, then **hydrochl**
crystals are made, e.g. copper sulfate, **sulfuric acid** is used. If **nitrate** crystals are
then **nitric** acid is used.

Metal + acid usually produces **flammable** hydrogen gas; therefore we tend to us
hydroxide/carbonate.

Quick check 4.2

1. Which ion is found in **all** acids?
2. (a) Give the colour change when hydrochloric acid is added to a solution of
   (b) What happens to the pH of a neutral solution when hydrochloric acid is
3. Which ion in aqueous solution causes alkalinity?
4. When hydrochloric acid is added to sodium h[illegible]xi[illegible] a reaction occurs to p
   (a) What type of reaction is this?
   (b) Name the salt produced
   (c) Describe how an i[illegible]r can be used to show when all the sodium hyd
   hydr[illegible]
5. Coppe[illegible]nate reacts with sulfuric acid to produce copper sulfate, water a
   observation that would be seen when copper carbonate is added to sulfuric a
6. Plan a method to produce a pure, dry sample of zinc nitrate from an acid and
7. Universal indicator is added to separate solutions of nitric acid, ammonia and
   Suggest the pH and colour of each solution.
   i) Nitric acid solution ii) Ammonia solution iii) Ammonium nit

4.3 Electrolysis

Background information

Liquids and solutions that can conduct electricity are called **electrolytes**.

Ionic compounds that are **molten** or **dissolved in water (aq)** are electrolytes because

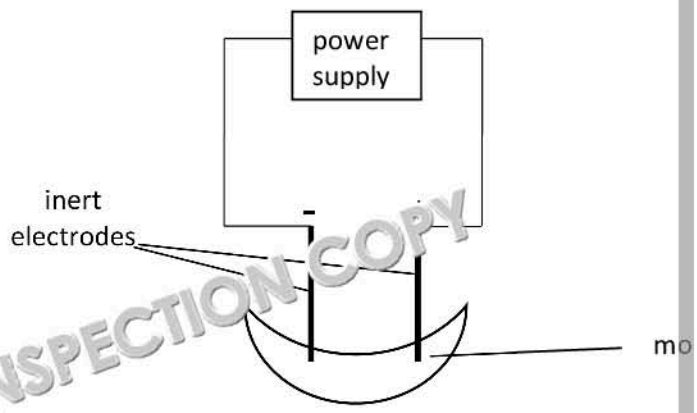
When we pass an electric current through an electrolyte:

- positive ions move to the cathode (negative electrode)
- negative ions move to the anode (positive electrode)

Ions are discharged at the electrodes, producing elements. The whole process is

Exam tip: Use the mnemonic PANIC: **positive (is) anode, negative is cathode**. The on the electrodes. Remember – opposites attract!

Worked example – molten ionic compounds



Electrolysis of molten lead bromide, $PbBr_2$

1. Split lead bromide into ions – Pb^{2+} and Br^-
2. Positive ion will move to the negative electrode (cathode – NIC)
3. Lead ions are discharged at the cathode, creating lead (Pb)
4. Negative ion will move to the positive electrode (anode – PA)
5. Bromide ions are discharged at the anode, creating bromine (Br_2)

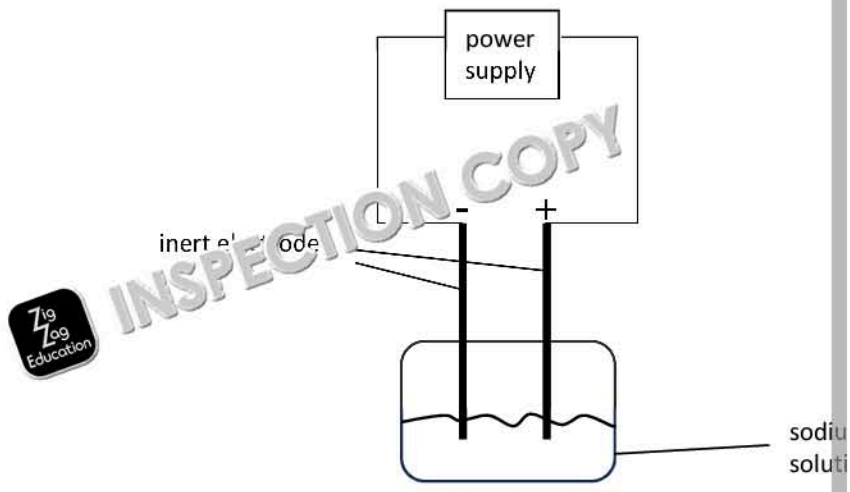
Exam tip: Note that at both electrodes the ions lose their charge and become ato pair up to form a bromine molecule as this gives bromine a stable electron config

In aqueous solutions things become slightly more complicated as water also splits ions can also be discharged.

At the cathode (**NIC**), we need to use the reactivity series of metal to decide which metal ion. If the metal is **more reactive** than hydrogen, H^+ is discharged and H_2 is

At the anode (PA), oxygen is produced unless the solution contains halide ions. In t

Worked example – aqueous ionic compounds



Electrolysis of aqueous sodium chloride, NaCl

1. Split sodium chloride into ions – Na^+ and Cl^-
2. Recall that solution means dissolved in water (aq); therefore the solution als – H^+ and OH^-
3. Positive ions (Na^+ and H^+) will move to the negative electrode (cathode – NIC
4. H+ ions will be discharged, creating H_2, as sodium is above hydrogen in the re
5. Negative ions (OH^- and Cl^-) will move to the positive electrode (anode – PA)
6. Chloride ions are discharged, creating Cl_2, as Cl^- is a halide ion

Exam tip: Key ideas – electrolysis

1. The **ionic** compound mu ve m lten or **dissolved in water (aq)** to ensure i
2. **Molten** compou n the ions of the **compound only**.
3. **Solu q)**, ntain both ions from the **compound** and ions from **water**.
4. Use P o remember which electrode is positive and which is negative.
5. Opposites attract.
6. **H_2** produced if **metal** is **above hydrogen** in reactivity series, otherwise **met**
7. **Halogen gas** produced if **halide ion** present, otherwise **oxygen gas** is produ

Electrolysis and extraction of metals

If a metal is above carbon in the reactivity series it can be extracted using electrol

Aluminium is manufactured by **electrolysis**. A molten mixture of **aluminium oxid** Cryolite is added to lower the melting point of aluminium oxide and reduce energ

The electrodes are made from **carbon**.

1. Split aluminium oxide into ions – Al^{3+} and O^{2-}
2. Positive ions, Al^{3+}, will move to the negati e ele de (cathode – NIC)
3. Al^{3+} ions will be discharged, creati A
4. Negative ions, O^{2-}, will m to t positive electrode (anode – PA)
5. Oxide ions are di d, reating O_2

The positive n electrode must be continually replaced. The oxygen gas prod reacts with the carbon electrode producing carbon dioxide gas. This increases the

Exam tip: Electrolysis is expensive due to the energy costs involved in heating th The costs of replacing the carbon electrodes must also be factored in if oxygen is carbon is a preferred method of extraction if the metal is less reactive than carb

Quick check 4.3

1. A student investigates the electrolysis of copper chloride solution.
   (a) Which gas is produced at the positive electrode?
   (b) Which ion is discharged at the negative electrode?

2. Name the products formed at each electrode when an aqueous solution of p undergoes electrolysis.
   i) Anode
   ii) Cathode

3. Name the products formed at each electrode when an aqueous solution of c undergoes electrolysis.
   i) Anode
   ii) Cathode

4. Aluminium is produced by the reduction of aluminium oxide.
   (a) What is meant by the term *reduction*?
   (b) Why is the anode continuously replaced during this reaction?

Checklist: Chemical changes

I can...
explain oxidation and reduction in terms of loss or gain of oxygen
identify substances which are oxidised or reduced
describe the reactions of metals with water (at room temperature)
describe the reactions of metals with dilute acids
explain how this reactivity is related to the metal's tendency to form positive ion
deduce an order of reactivity of metal based on experimental results (displacem
evaluate the best method of extraction for a metal from information given
predict the products of neutralisation reactions when given the reactants
describe how to make a pure, dry sample of a soluble salt for information given
describe the use of an indicator to measure the pH of a solution
use the pH scale to identify acidic or alkaline solutions
predict the products of the electrolysis of molten ionic compounds
predict the products of the electrolysis of aqueous ionic compounds
explain why a mixture is used as an electrolyte in the extraction of aluminium
explain why the positive electrode must be continuously replaced in the extracti

Chapter 5: Energy cha

5.1 Exothermic and endothermic re

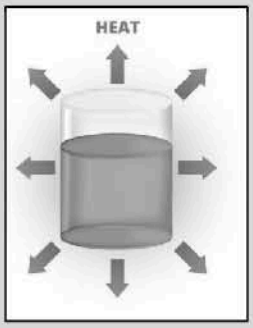
Background information

Conservation of energy

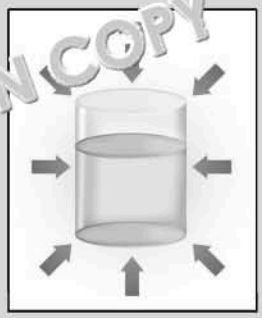
Energy cannot be created or destro[illegible] he[illegible]ove, energy is **conserved** during a c[illegible]

Exothermic reactions t[illegible] **ergy to** the surroundings. The temperature of th[illegible]
Examples c[illegible] reactions include combustion, oxidation and neutralisatio[illegible]

Self-heating cans and hand warmers use exothermic reactions.



Endothermic reactions **take in energy from** the surroundings. The temperature o[illegible]
An example of an endothermic reaction is thermal decomposition. Sports injury pa[illegible]



Exam tip: Exothermic energy **ex**its. **En**dothermic energy **en**ters.

Investigating the variables that affect temperature changes in reac[illegible]

This procedure is identical for the following reactions which we have met in previ[illegible]

- acid plus metal
- acid plus carbonates
- neutralisations
- displacement reactions of metals

It is also valid for other reactions that involve tem[illegible]e changes.

Procedure

1. Place a polystyrene [illegible] glass beaker to make it stable.
2. Add 25 [illegible] of [illegible]tal solution.
3. Record [illegible]mperature.
4. Add ..... [illegible] ..... g of the metal/carbonate/alkali.
5. Record highest/lowest temperature reached.
6. Repeat

If the temperature increases – exothermic reaction.
If the temperature decreases – endothermic reaction.

Variables

To ensure an experiment is a fair test we must control all the variables except the

- ★ **Control variables**: All the variables we keep the same during the experiment,
- ★ **Dependent variable**: The variable we are measuring – temperature change.
- ★ **Independent variable**: The variable we are changing.

Example of variables that you might change during different experiments:

- mass of metal/carbonate added to acid, e.g. 1 g or 2 g
- surface area of the carbonate, e.g. powder or lumps
- volume of alkali added to the acid, e.g. 25 cm^3 or 10 cm^3
- type of acid, e.g. hydrochloric acid or sulfuric acid
- type of metal, e.g. zinc or iron
- type of metal solution, e.g. copper sulfate or zinc nitrate

In all experiments we are trying to determine how changing the independent variable affects the dependent variable (temperature).

Worked examples

Temperature changes in reacting solutions – displacement reactions

Four different metals were added to separate solutions of copper sulfate. The temperature change was recorded.

metal	magnesium	zinc	iron
temperature change (°C)	+38	+24	+11

a) How do the results show the reaction with three of the metals is **exothermic**?

Answer: The **temperature increases** as energy is transferred to the surroundings.

b) How can you ensure the experiment is a **fair test**?

Answer: Add the **same mass** of each metal and use the **same amount/volume**

Temperature changes in reacting solutions – thermal decomposition

Magnesium carbonate decomposes to produce magnesium oxide and carbon dioxide. The temperature decreases from 27 °C to 18 °C. Carbon dioxide is released as a gas.

a) What is the temperature change in this reaction?

Answer: 27 – 18 = 9
Temperature decreases by **-9 °C**

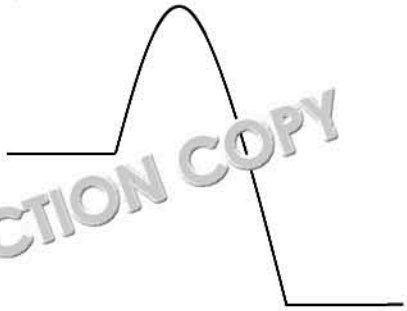
b) What type of reaction is happening when the temperature decreases?

Answer: Endothermic

Exam tip: You may be given a procedure on the exam and asked to improve the experiment. Usually, the examiner is looking for you to understand about variables: dependent and independent variables for the procedure given. Make sure the control amount/volume/mass. Next, decide the most appropriate way to change the independent information given.

Quick check 5.1

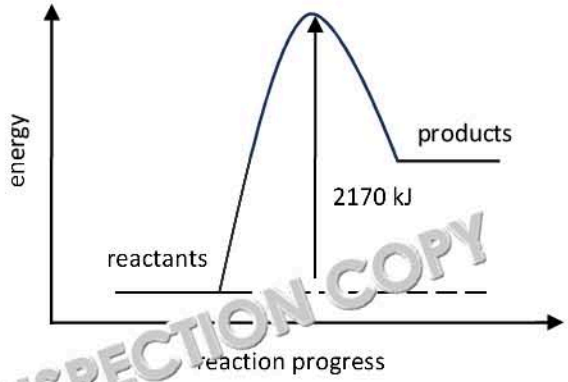
1. The diagram shows a reaction profile for a combustion reaction.



What type of reaction is shown in the diagram?

2. State the missing word in the sentence below.
In an exothermic reaction the energy of the reactants is **(a)** *than the energy o*

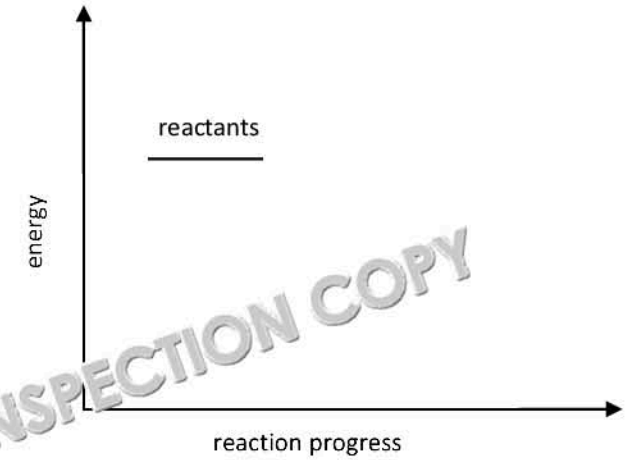
3. The diagram shows the energy change for a decomposition reaction.



What the energy value 2170 kJ represent?

4. The reaction between an acid and an alkali is exothermic.

Copy and complete the energy profile diagram for the reaction. Label the ac
energy change of the reaction.



5.2 Reaction profiles

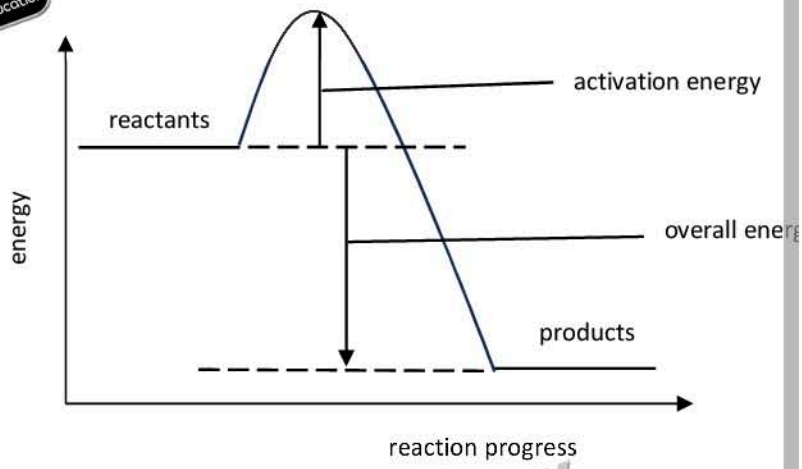
Background information

Reactions occur when particles collide with sufficient energy.

Activation energy – the minimum amount of energy that particles must have to r

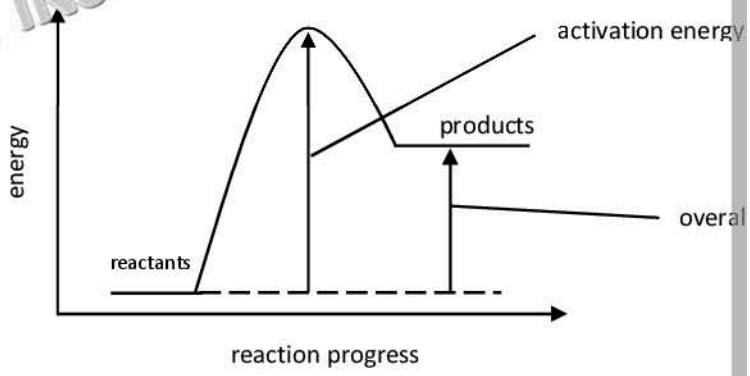
Reaction profiles are used to show the relative energies of the reactants and prod
the overall energy change of a reaction

Exothermic reactions



Note that the energy of the products is **less** than the energy of the reactants. Ene
reaction and there is a rise in temperature of the surroundings.

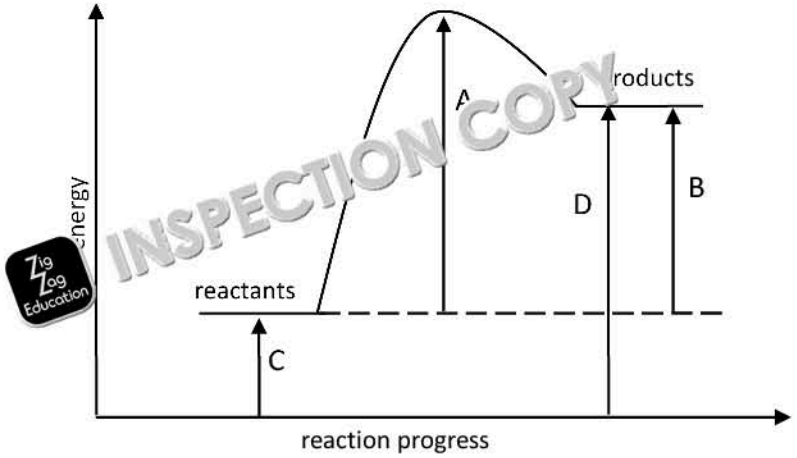
Endothermic reactions



Note that the energy of the products is **more** than the energy of the reactants. En
reaction and there is a decrease in temperature of the surroundings.

Worked example

The diagram shows the reaction profile for an endothermic reaction.



a) Which letter represents the activation energy for the reaction?

Answer: A – activation energy is the energy from the reactants to the highest

b) Which letter represents the overall energy change for the reaction?

Answer: B – overall energy change is the energy difference between reactant

Quick check 5.2

1. In a reaction between an acid and a [illegible] the temperature increases. What when the temperature increases?
2. A student investigates the change in temperature when different masses of m sulfate [illegible]on.
   Plan a suitable method for this investigation.
3. Name an item that uses:
   (a) an exothermic reaction (b) an endothermic
4. A student wants to investigate the reactivity of different metals with nitric ac
   (a) Name **two** variables that must be kept constant. (b) What is the indep
5. A metal carbonate takes in energy to break down (decompose).
   What type of reaction takes in energy from the surroundings?

Checklist: Energy change

I can...
distinguish between exothermic and endothermic reactions based on temperatu the surroundings
evaluate the uses and applications of exothermic and endothermic reactions
investigate [illegible] that affect temperature changes in reacting solutions
draw simple energy profile diagrams for exothermic and endothermic reactions
show the activation energy and overall energy change on a reaction profile diagr exothermic and endothermic reactions
use reaction profile diagrams to identify reactions as exothermic or endothermic
explain that activation energy is the energy needed for a reaction to occur

Answers

Chapter 1: Atomic structure and the pe

Quick check 1.1: Atoms, elements, compounds and mixtures

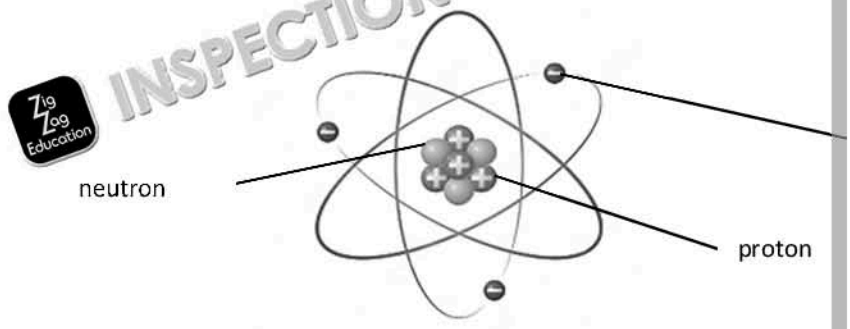
1. (a) chlorine + potassium bromide → potassium chlo... + bromine
   (b) hydrochloric acid + sodium hydroxide → so...um... ...ide + water
2. (a) **2**Na + Cl_2 → **2**NaCl
   (b) N_2 + **3**H_2 → **2**NH_3
   (c) **2**Mg + O_2 → **2**Mg...
3.

name	formula
water	H_2O
methane	**CH_4**
ammonia	**NH_3**
hydrochloric acid	**HCl**
sulfuric acid	**H_2SO_4**
carbon dioxide	**CO_2**

4. B
5. Place the mixture of salt and sand in a beaker, add water. The salt will dissolve. Filte... filter paper. The sand will remain in the filter paper. The salt and water will pass thr... conical flask.

Quick check 1.2: The periodic table

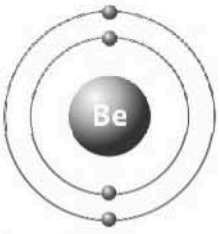
1. Two conclusions:
   - atoms are mostly empty space
   - the mass of an atom is concentrated at its c...ntre
2.



3.

48 **Ti** titanium 22	28 **Si** silicon 14	16 **O** Oxygen 8	20 **Ne** neon 10

4. **Mg** – protons = 12, electrons = 12, neutrons 24 – 12 = 12
   He – protons = 2, electrons = 2, neutrons 4 – 2 = 2
   Al – protons = 13, electrons = 13, neutrons 27 – 13 = 14
   S – protons = 16, electrons = 16, neutrons 32 – 16 = 16
   Ca – protons = 20, electrons = 20, neutrons 40 – 20 = 20
5. (a) Number of el... **2**
   N... of ...tons: **2**
   Na... element: **helium**
   (b) Number of electrons: **3**
   Number of neutrons: **4**
   Name of element: **lithium**
6. Isotopes: atoms of the same element with the same number of protons but different

7. 

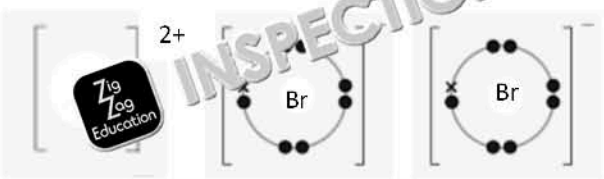
8. Be: 2,2
   P: 2,8,5
   K: 2,8,8,1
9. Atomic weight
10. gaps
11. Similari both have three electron shells. / They both have an incomplete oute
    Differen otassium has one electron in its outer shell and bromine has seven.

Quick check 1.3: Groups, properties and trends

1. Alkali metals
2. **2**Li + **2**H_2O → **2**LiOH + H_2
3. Lithium fizzes steadily (H_2 gas produced); slowly becomes smaller and eventually dis
   addition of universal indicator)
   Sodium fizzes rapidly and forms a ball; quickly becomes smaller and disappears
   addition of universal indicator)
4. (a) sodium + chlorine → sodium chloride
   (b) Lithium atom loses one electron and forms a positive ion. Chlorine atom gains o
5. (a) Number of protons: 17
   Number of neutrons: 35 – 17 = 18
   (b) Isotopes
6. Metals and non-metals
7. Iodine. Atoms gain one outer electron to for a egat ly charged ion. Going dow
   outer electron is further from nucle or attraction between nucleus and oute
   electron is harder to gain.
8. Melting points, boili t. nd molecular mass all increase going down the group
9. (a) ch + p m iodide → **iodine + potassium chloride**
   (b) Di ent
10. (a) No gases
    (b) 2,8
    (c) Neon is inert as it has a full outer shell of electrons and a stable electron confi

Chapter 2: Structure and bondi

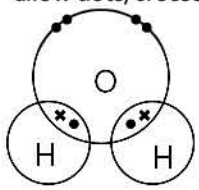
Quick check 2.1: Ionic bonding

1. Ionic
2. Beryllium oxide forms a **giant** structure of oppositely charged ions (ionic lattice) held
   attraction. It has a high **melting** point and conducts electricity when in **molten** state
3. Ca 2,8,8,**2** S 2,8,**6** Ca^{2+} 2,8,**8** S^{2-} 2,8,**8**
   Ca **atom** has **2 electrons** on its outer shell. The Ca at ns ers these **2** outer elec
   Ca^{2+} **ion**. The S **atom accepts** these **2 electro** s a d h mes a S^{2-} **ion**. Both ions no
   and are stable. This is called **ionic b** in
4. 

5. Electrostatic attraction
6. (a) Ions are fixed
   (b) Melt it / Dissolve it
   (c) Ions are mobile

Quick check 2.2: Covalent bonding

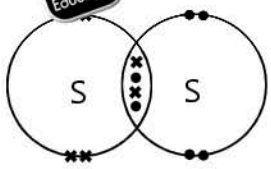
1. (covalent) bonds
2. (a) allow dots/crosses either way around



(b) No information about sharing of electrons

(c) Does not show how the atoms are arranged in 3D space

3. $C_1H_2Cl_2$ = CH_2Cl_2
4. (a) C
5. (b)

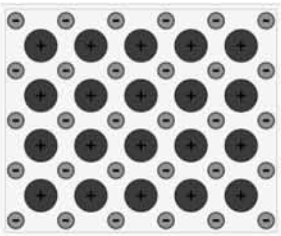
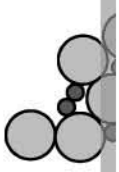


(c) The boiling point of sulfur is **low**. This is because sulfur contains **weak** forces b

6. a fullerene
7. Covalent
8. The atoms in a polymer molecule are held together by **covalent** bonds.
   The polymer chains are held together by **intermolecular** forces.
   When a polymer melts or boils, the **intermolecular forces** are broken.
8. (a) In graphite, each carbon atom forms bonds with **four** other carbon atoms. Gra
   delocalised electrons.
   (b) Soft/slippery, low melting point
   (c) It is soft/slippery and the layers can slide over each other.
9. The structure of a fullerene is based on **hexagons**.
   The fullerene molecule is made from **carbon**.
   One use of the fullerene molecule is **nanotubes / nanotechnology / materials / elec**

Quick check 2.3: Metallic bonding

1.

Metal	Alloy
	
Pure metals are soft because their atoms are arranged in layers and can slide easily when a force is applied.	Alloys are harder and strong of atoms of different sizes. T disrupted and require a large make them slide.

2. Copper alloy is a mixture of copper and another element. The atoms of each eleme the regular arrangement of copper atoms and makes it harder for the atoms to slide

Quick check 2.4: Bonding, structure and properties

1. + $H_2O(l)$ ⟶ NaOH(**aq**)
2. Condensing/condensation
3. Solid
4. (a) State symbol
   (b) Aqueous

Chapter 3: Quantitative chemis

Quick check 3.1: Balanced equations and conservation of mass

1. **2**$K + 2H_2O \rightarrow 2KOH + H_2$
2. 4.1 + X = 9.6; rearrange this to 9.6 – 4.1 = X
   Therefore X = 5.5 g
3. (a) Two products are formed and the sum of both prod[illegible] masses is equal to the
   (b) 100 – 56 = 44 g

Quick check 3.2: Relative ato[illegible] (A_r)

1. $\frac{(15 \times 15) + (16 \times 50) + [illegible]}{10} = 16.2$

Quick check 3.3: Relative formula mass (M_r)

1. 24 + 32 + (4 × 16) = 120
2. (a) 123.5 – 12 – (3 × 16) = 63.5
   (b) Cu, copper
3. 16 / 40 × 100 = 40 %

Quick check 3.4: Chemical measurements

1. (a) Mean = $\frac{54.3 + 55.4 + 56.1}{3}$ = 55.3 g
   (b) Range = 56.1 – 54.3 = 1.8
   Uncertainty = 1.8 / 2 = 0.9 g 55.3 g ± 0.9 g

Quick check 3.5: Concentration of solutions

1. 200 / 50 = 4; 6 / 4 = 1.5 g

Chapter 4: Chemical change

Quick che[illegible] Reactivity of metals

1. (a) Nitric acid
   (b) Salt
2. Magnesium

Quick check 4.2: Reactivity of acids

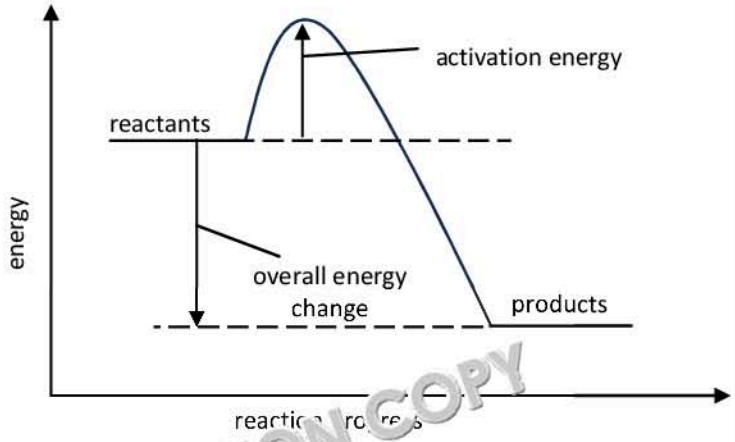
1. H^+
2. (a) Green to red
   (b) Decreases
3. OH^-
4. (a) Neutralisation
   (b) Sodium chloride
   (c) Add indicator to sodium hydroxide and it will turn dark blue. When adding the
   change from dark blue to light blue to green (neutra[illegible]ve add more acid, the
   orange/red, indicating the solution is now a[illegible]
5. The solid will remain at the bottom of t[illegible]er as it is in excess.
   Bubbles of carbon dioxide gas [illegible]
6. Dissolve powdered zi[illegible]ric acid.
   Stir the r[illegible]on [illegible]and continue adding zinc oxide until it is in excess.
   Filter t[illegible]re to remove excess zinc oxide.
   Heat ge[illegible]an evaporating dish until crystals begin to appear.
   Leave solution in a warm place for a few days.
   Dry the zinc nitrate crystals between two sheets of filter paper.
7. i) pH 2–3, red/orange
   ii) pH 11–14, blue
   iii) pH 5–7, yellow/green

Quick check 4.3: Electrolysis

1. (a) Chlorine
   (b) Copper/Cu^{2+}
2. i) Bromine
   ii) Hydrogen
3. i) Oxygen
   ii) Copper
4. (a) Loss of oxygen / Gain of electrons
   (b) Oxygen produced during the reaction reacts with the carbon anode producing

Chapter 5: Energy changes

Quick check 5.1: Exothermic and endothermic reactions

1. Exothermic
2. Higher/more
3. Activation energy
4. 

Quick check 5.2: Reaction profiles

1. Exothe
2. Place a polystyrene cup inside a glass beaker to make it stable.
   Add 25 cm^3 of copper sulfate solution.
   Record the temperature.
   Add 1 g of the magnesium metal.
   Record highest/lowest temperature reached.
   Repeat with different masses of magnesium.
3. (a) Hand warmer / Self-heating can
   (b) Sports injury pack
   (There are alternatives but these are the examples listed in the specification)
4. (a) Volume of nitric acid, mass of the metals
   (b) Type of metal
5. Endothermic

Appendix: Write-on quick chec

Quick check 1.1

1. (a) Chlorine reacts with potassium bromide to produce bromine and potass equation for this reaction.

   .... + ... → .... + ...

   (b) Hydrochloric acid reacts with sodium hydroxide to produce sodium chlo equation for this reaction.

   .... + ... → .... + ...

2. Balance the chemical equations for the following reactions:

   (a) ......Na + ...Cl_2 → ......NaCl

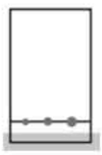
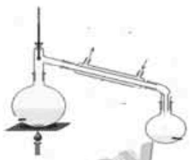
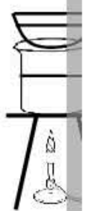
   (b) ......N_2 + ...H_2 → ......NH_3

   (c) ......Mg + ...O_2 → ......MgO

3. Use the Internet to research the formula for each compound given in the tab The first one has been done for you.

name	formula
water	H_2O
methane	
ammonia	
hydrochloric acid	
sulfuric acid	
carbon dioxide	

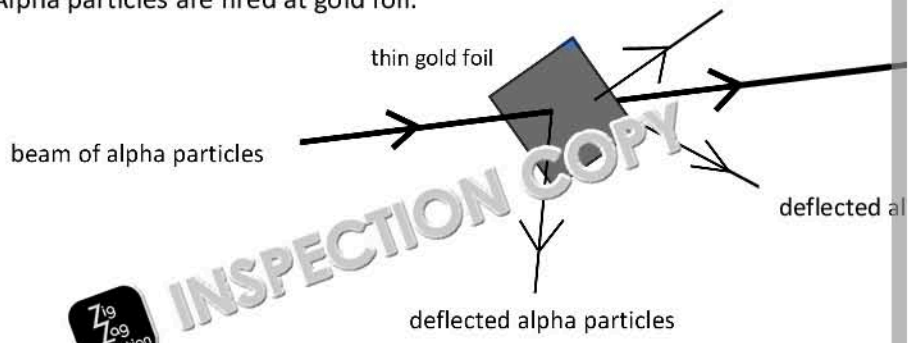
4. Which apparatus can be used to separate different food colourings? Circle o

A  B  C  D 

5. Describe how you would separate a mixture of salt and sand. Use diagrams t

Quick check 1.2

1. Alpha particles are fired at gold foil.

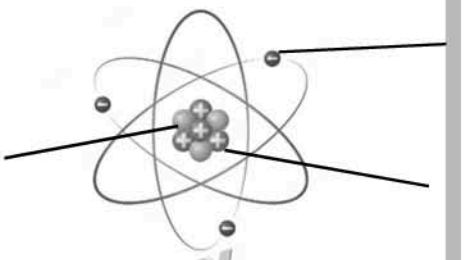


Write down two conclusions that can be made from these results.

1. .....

2. .....

2. Label the subatomic particles on the diagram on the right.



3. Complete the element names and symbols th[illegible] re missing from the exampl[illegible] to help you.

48	[illegible]	16	20
22	14	8	10

4. How many protons, electrons and neutrons do atoms of the following eleme[illegible]

24 **Mg** magnesium 12	4 **He** helium 2	27 **Al** aluminium 13	su
Protons	.....	.....	
Electrons	.....	.....	
Neutrons	.....	.....	

5. Figures 1 and 2 show a R[illegible] model of an atom of two different elements in t[illegible]

Figure 1

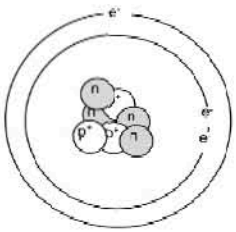


Figure 2

(a) What is the number of electrons and neutrons in **Figure 1**'s atom?

Number of electrons: ..... Number of neutrons: .....

(b) What is the number of electrons and neutrons in **Figure 2**'s atom?

Number of electrons: ..... Number of neutrons: .....

6. What is meant by isotopes? (You should refer to subatomic particles in your

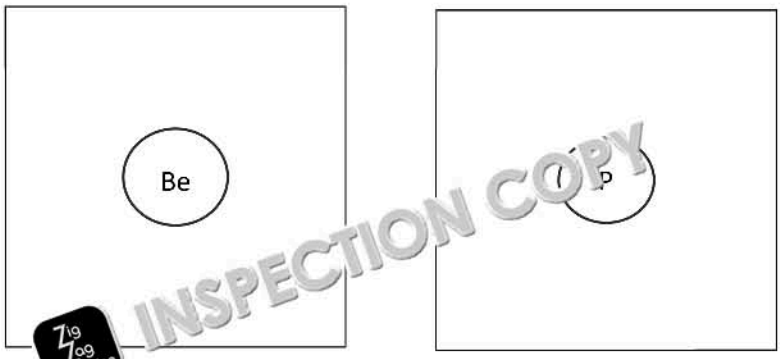
.....

.....

.....

7. Draw the electron configuration diagrams for an atom of each of the followi

9	31	39
Be	**P**	**K**
beryllium	phosphorus	potassiu
4	15	19



8. Write down each electron configuration in number format.

Be ..... P .....

9. What order did scientists use to arrange elements in early periodic tables?

.....

10. Complete the sentence.

Mendeleev overcame the problem associated with the early periodic table by for elements that had not yet been discovered.

11. State one similarity and one difference in the electronic structure of potassiu

Similarity: .....

.....

Differe .....

.....

Quick check 1.3

1. Lithium, sodium and potassium are group 1 elements. What is the name give

.....

2. Lithium reacts with water to produce lithium hyd[illegible]e and hydrogen. Balan this reaction.

..... Li + ... H_2O → ..... LiOH +

3. Group [illegible]als react with water. Compare what is seen when lithium rea reacts [illegible]ater.

.....

.....

.....

4. Sodium reacts with chlorine to produce sodium chloride.

(a) Write a word equation for this reaction.

..... + ...→ .....

(b) Describe what happens when a lithium atom reacts with a chlorine atom transfer in your answer.

.....

.....

5. An ato[illegible]lorine is represented as $^{35}_{17}Cl$

(a) Determine the number of protons and the number of neutrons in one at

Number of protons: .....

Number of neutrons: .....

As well as $^{35}_{17}Cl$, chlorine atoms can be of the form $^{37}_{17}Cl$

(b) What is the name given to these different atoms of chlorine?

.....

6. Which **types** of element react when M[illegible] ac[illegible] with Cl_2?

.....

7. The halogens Cl_2, Br_2 and I_2 all react with silver. Which halogen is the **least** re

.....

Explain your answer.

.....

.....

8. Give two trends in the **ph**[illegible]**al** p[illegible]perties of the halogens.

1. .....

2. .....

9. Chlorine reacts with potassium iodide solution.

(a) Complete the word equation.

chlorine + potassium iodide → ..... + ...

(b) What is the name of this type of reaction?

.....

10. (a) What is the name of the group that contains neon?

.....

(b) What is the electronic stru[illegible] o[illegible] neon atom?

.....

(c) Ex[illegible]hy **no** products are formed when neon is burned in oxygen.

.....

.....

Quick check 2.1

1. The diagram shows a dot and cross diagram for potassium fluoride.

What type of bonding is present in potassium fluoride?

.....

2. Complete the paragraph below. Use words from the box.

strong	**melting**	**molten**
attraction	**mobile**	**giant**

Beryllium oxide forms a _____ structure of oppositely charged io

_____ forces of _____. It has a high _____

when in _____ state due to _____ ions.

3. Describe how a calcium atom and a sulfur atom will react to form calcium su structure of each atom and the electronic structure and charges of any ions f

.....

.....

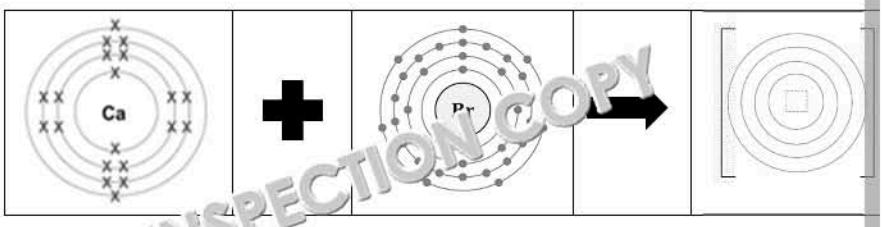
.....

.....

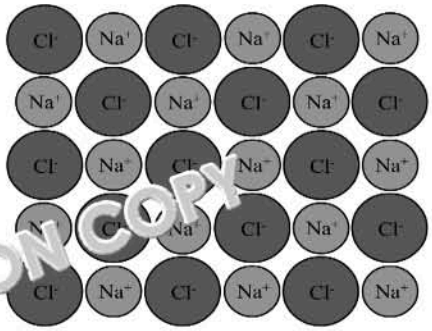
.....

.....

4. Calcium reacts with bromine to produce calcium bromide. Complete the dot this reaction.



5. The diagram shows the structure of an ionic lattice. Which force holds the par



.....

6. Solid ionic compounds do not conduct electricity.

(a) Explain why.

.....

.....

(b) How can you make an ionic compound conduct electricity?

.....

.....

(c) Explain how this makes the ionic compound conduct electricity.

.....

.....

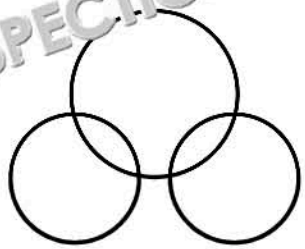
Quick check 2.2

1. The diagram shows a model of a nitrogen molecule.

What do the lines between the atoms represent?

.....

2. (a) Complete the dot and cross diagram for a molecule of water. Show only



(b) The diagram shows a ball and stick model of a water molecule.

Give **one** limitation of using the ball and stick model of water over a dot and cross model.

.....

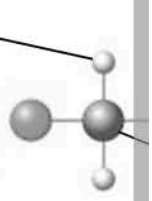
.....

(c) Give **one** limitation of using the dot and cross model of water over the b

.....

.....

3. Give the molecular formula of the molecule on the right.



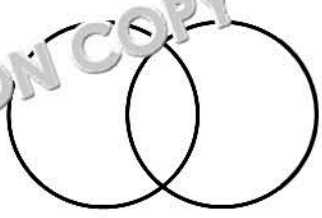
.....

4. Sulfur is below oxygen in the periodic table.

(a) What type of bonding is present in molecules of sulfur?

.....

(b) Deduce the dot and cross diagram for a molecule of sulfur, S_2. Show only



(c) Complete the sentences by crossing out the incorrect words in bold.

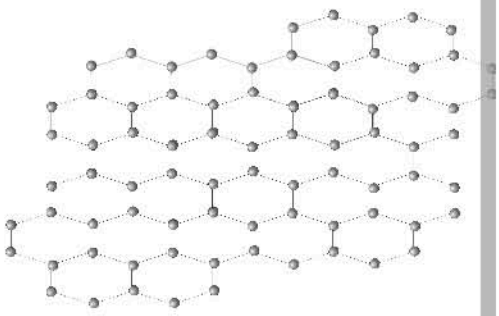
*The boiling point of sulfur is **low/high**. This is because sulfur contains **we** between its molecules.*

5. The diagram shows the structure of C_{60}.

Circle the correct answer. C_{60} is

giant ionic	graphite	metallic	a full

6. Graphene is a single layer of graphite.



What type of bonding is present between the carbon atoms in graphene?

.....

7. Complete the sentences.

The atoms in a polymer molecule are held together by _____

The polymer chains are held together by _____ *forces*

When a polymer melts or boils, the _____ *are broken*

8. Complete the sentences.

(a) *In graphite, each carbon atom forms bonds with* _____ *other carbon atoms*

electricity due to _____.

(b) Give **two other** physical properties of graphite.

1. .....

2. .....

(c) Suggest why graphite can be used as a lubricant

.....

.....

9. Complete the sentences.

The structure of a fullerene is based on _____

The fullerene molecule is made from atoms of _____

One use of the fullerene molecule is _____.

Quick check 2.3

1. Explain how alloying a metal improves its properties. Draw diagrams to help

Metal	
.....	.....
.....	.....
.....	.....
.....	.....

2. Explain why a copper alloy is harder than pure copper metal.

.....

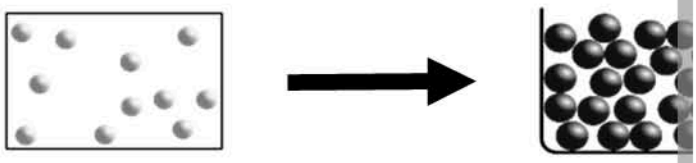
.....

Quick check 2.4

1. Complete the chemical equation using state symbols. Sodium metal reacts with water to form a solution of sodium hydroxide and hydrogen gas.

Na(....) + H_2O(....) → NaOH(....)

2. A gaseous molecule is cooled to form a liquid during an experiment. Name the process.



.....

3. The melting point of lithium is 180 °C. What is the state of lithium at 150 °C?

.....

4. (a) What is the general term used for the letter(s) given in brackets in an equation for a compound?

.....

(b) What does (aq) mean?

.....

Quick check 3.1

1. Potassium reacts with water to produce potassium hydroxide and hydrogen.

   Balance the equation for this reaction.

.....K	+	2'H2(	→	2KOH

2. 4.1 g of potassium r... w... fluorine to produce 9.6 g of potassium fluorid... that re...

   .....

   .....

3. Solid calcium carbonate ($CaCO_3$) decomposes to produce calcium oxide solid ... (CO_2). 100 g of calcium carbonate was heated during an experiment and 56 g ...

   (a) Why is the mass of calcium oxide **less** than the mass of calcium carbona...

   .....

   .....

   (b) What mass of carbon dioxide is formed?

   .....

Quick chec... 3.2

1. The mass numbers and percentage abundance of three isotopes of oxygen a...

mass number	percentage abund...
15	15
16	50
17	35

   Calculate the relative atomic mass of oxygen.

   .....

   .....

   .....

Quick check 3.3

1. Calculate the relative formula mass of $MgSO_4$. A_r: Mg = 24, S = 32, O = 1

.....

.....

.....

2. The relative formula mass of **X** [illegible] Relative atomic masses A

   (a) Calculate the [illegible]omic mass of **X**

   .....

   .....

   .....

   (b) Name X .....

3. Magnesium reacts with oxygen to produce magnesium oxide, MgO. Calculate oxygen in magnesium oxide (A_r: O = 16, M_r: MgO = 40).

.....

.....

.....

Quick check 3.4

1. **Table 1** shows the results of th[illegible]ments when solid calcium carbonate calcium oxide solid and [illegible]n dioxide gas.

	mass of calcium oxide produced (g)	
1	2	3
54.3	55.4	56.1

Table 1

   (a) Complete the table to show the mean value.

   .....

   .....

   (b) Calculate the uncertainty in the range of measurements.

   .....

   .....

   .....

Quick check 3.5

1. A copp[illegible]ride solution contains 6 g of copper chloride in 200 cm^3 of solut copper [illegible]de in 50 cm^3 of this solution.

.....

.....

.....

Quick check 4.1

1. (a) An acid reacts with zinc to produce zinc nitrate and hydrogen. Which ac
zinc nitrate?

.....

(b) What type of substance is zinc nitrat ?

.....

2. Calciu an unknown metal X were added to separate solutions of zinc
was re

metal	tempera
calcium	
X	

Suggest the name of metal X: .....

Quick check 4.2

1. Which ion is found in **all** acids?

.....

2. (a) Give the colour change when hydrochloric acid is added to a solution of

.....

(b) What happens to the pH of neutral solution when hydrochloric acid is

.....

3. Which aqueous solution causes alkalinity?

.....

4. When hydrochloric acid is added to sodium hydroxide, a reaction occurs to p

(a) What type of reaction is this?

.....

(b) Name the salt produced.

.....

(c) Describe how an indicator can be used to show when all the sodium hyd hydrochloric acid.

.....

.....

.....

5. Copper carbonate reacts with sulfuric acid to produce copper sulfate, water a observation that would be seen when copper carbonate is added to sulfuric a

1. .....

.....

2. .....

.....

6. Plan a method to produce a pure, dry sample of zinc nitrate from an acid and

.....

.....

.....

.....

7. Univer ca or is added to separate solutions of nitric acid, ammonia and Sugges pH and colour of each solution.

i) Nitric acid solution: .....

ii) Ammonia solution: .....

iii) Ammonium nitrate solution: .....

Quick check 4.3

1. A student investigates the electrolysis of copper chloride solution.

   (a) Which gas is produced at the positive electrode?

   .....

   (b) Which ion is discharged at the negative electrode?

   .....

2. Name the products formed at each electrode when an aqueous solution of p undergoes electrolysis.

   i) Anode: .....

   ii) Cathode: .....

3. Name the products formed at each electrode when an aqueous solution of c undergoes electrolysis.

   i) Anode: .....

   ii) Cathode: .....

4. Aluminium is produced by the electrolysis of aluminium oxide.

   (a) What is meant by the term *reduction*?

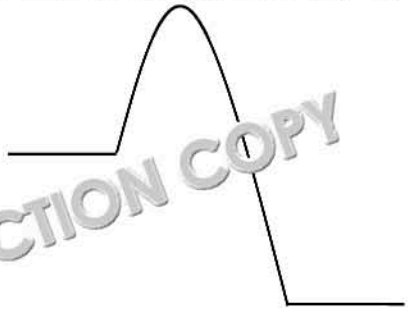
   .....

   (b) Why is the anode continuously replaced during this reaction?

   .....

Quick check 5.1

1. The diagram shows a reaction profile for a combustion reaction.



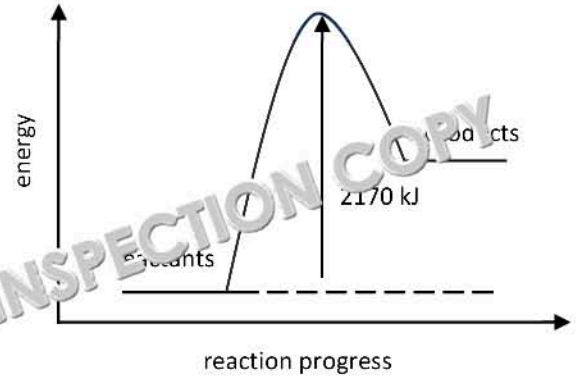
What type of reaction is shown in the diagram?

.....

2. Complete the sentence.

In an exothermic reaction the energy of the reactants is _____ *the products.*

3. The diagram shows the energy change for a decomposition reaction.

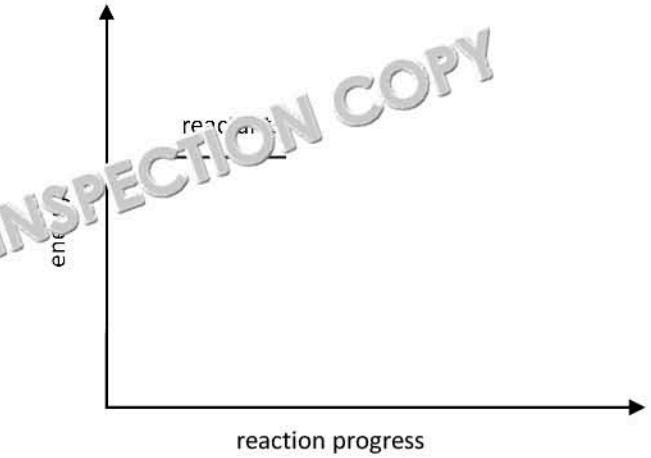


What does the energy value 2170 kJ represent?

.....

4. The reaction between an acid and an alkali is exothermic.

Complete the energy profile diagram for the reaction. Label the activation energy and the energy change of the reaction.



Quick check 5.2

1. In a reaction between an acid and a metal, the temperature increases. What when the temperature increases?

   .....

2. A student investigated the change in temperature when different masses of m sulfate solution.

   Plan a six-step method for this investigation.

   1. .....
   2. .....
   3. .....
   4. .....
   5. .....
   6. .....

3. Name an item that uses:

   (a) an exothermic reaction

   .....

   (b) an endothermic reaction

   .....

4. A student wants to investigate the reactivity of different metals with nitric ac

   (a) Name **two** variables that must be kept constant.

   1. .....
   2. .....

   (b) What is the independent variable in this reaction?

   .....

5. A metal carbonate takes in energy to break down (decompose).

   What type of reaction takes in energy from the surroundings?

   .....