



Chemistry Paper 1 Knowledge and Exam Practice

Monday 22nd May 2023

Foundation

Topics – Atomic Structure and the periodic table, Structures and Bonding, Chemical Changes and Energy changes

How to use this booklet



Rehearse the content using to knowledge organiser pages.

You can look-cover-say-check.

Make flash cards for content you find more difficult.

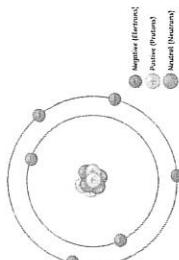
Then complete the exam practice questions.

Atomic Structure and the Periodic Table – Foundation and Higher

Atoms

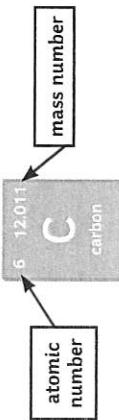
Contained in the nucleus are the protons and neutrons. Moving around the nucleus are the electron shells. They are negatively charged.

Particle	Relative Mass	Charge
proton	1	+1
neutron	1	0
electron	Very small	-1



Overall, atoms have no charge; they have the same number of protons as electrons. An ion is a charged particle - it does not have an equal number of protons to electrons.

Atomic Number and Mass Number



Elements are made of atoms with the same atomic number. Atoms can be represented as symbols.

N = nitrogen F = fluorine Zn = zinc Ca = calcium

Isotope	Protons	Electrons	Neutrons
¹ H	1	1	1 - 1 = 0
² H	1	1	2 - 1 = 1
³ H	1	1	3 - 1 = 2

Isotopes – an isotope is an element with the **same number of protons** but a **different number of neutrons**. They have the **same atomic number**, but different mass number.

Compounds – a compound is when two or more elements are chemically joined. Examples of compounds are carbon dioxide and magnesium oxide. Some examples of formulas are CO_2 , NaCl , HCl , H_2O , Na_2SO_4 . They are held together by chemical bonds and are difficult to separate.

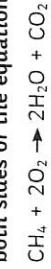
Equations and Maths

To calculate the relative atomic mass, use the following equation:

$$\text{relative atomic mass (A_r)} = \frac{\text{sum of (isotope abundance} \times \text{isotope mass number})}{\text{sum of abundances of all isotopes}}$$

Balancing Symbol Equations

There must be the same number of atoms on both sides of the equation:



$$\text{C} = 1$$

$$\text{O} = 4$$

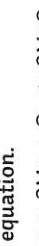
$$\text{H} = 4$$

Chemical Equations

A chemical reaction can be shown by using a word equation.

e.g. magnesium + oxygen \rightarrow magnesium oxide
On the left-hand side are the reactants, and the right-hand side are the products.

They can also be shown by a symbol equation.



Mixtures, Chromatography and Separation

Mixtures – in a mixture there are no chemical bonds, so the elements are easy to separate. Examples of mixtures are air and salt water.



Chromatography – to separate out mixtures.

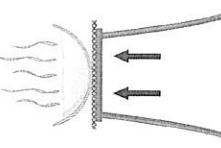
piece of wood

paper

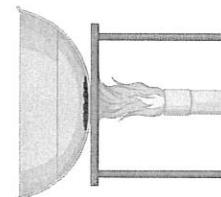
beaker

water

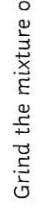
start



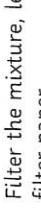
Filtration – to separate solids from liquids.



Evaporation – to separate a soluble salt from a solution; a quick way of separating out the salt.



Crystallisation – to separate a soluble salt from a solution; a slower method of separating out salt.



Separating out salt from rock salt:

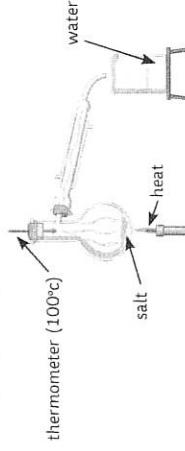
1. Grind the mixture of rock salt.
2. Add water and stir.
3. Filter the mixture, leaving the sand in the filter paper
4. Evaporate the water from the salt, leaving the crystals.

Atomic Structure and the Periodic Table – Foundation and Higher

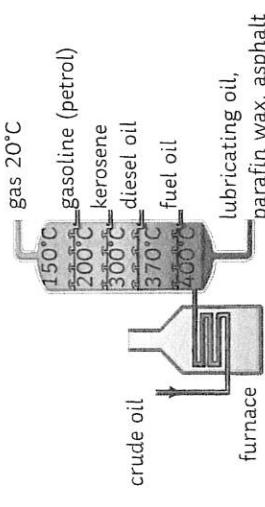
Distillation

To separate out mixtures of liquids.

- Simple distillation** – separating a liquid from a solution.



- Fractional distillation** – separating out a mixture of liquids. Fractional distillation can be used to separate out crude oil into fractions.



History of the Atom

Scientist	Time	Discovery
John Dalton	start of 19 th century	Atoms were first described as solid spheres.
JJ Thomson	1897	Plum pudding model – the atom is a ball of charge with electrons scattered.
Ernest Rutherford	1909	Alpha scattering experiment – mass concentrated at the centre; the nucleus is charged. Most of the mass is in the nucleus. Most atoms are empty space.
Niels Bohr	around 1911	Electrons are in shells orbiting the nucleus.
James Chadwick	around 1940	Discovered that there are neutrons in the nucleus.

Development of the Periodic Table

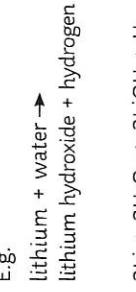
In the early 1800s, elements were arranged by atomic mass. The periodic table was not complete because some of the elements had not been found. Some elements were put in the wrong group.

Dimitri Mendeleev (1869) left gaps in the periodic table. He put them in order of **atomic mass**.

The gaps show that he believed there was some undiscovered elements. He was right! Once found, they fitted in the pattern.

They react with water and produce hydrogen.

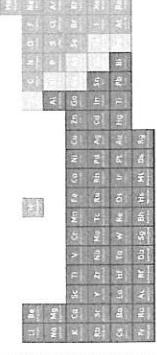
E.g.



The Modern Periodic Table

Elements are in order of **atomic mass/proton number**. It shows where the metals and non-metals are. **Metals** are on the **left** and **non-metals** on the **right**. The **columns** show the **groups**. The **group number** shows the number of **electrons** in the **outer shell**. The rows are **periods** – each period shows another full shell of electrons.

The periodic table can be used to predict the reactivity of elements.



Group 7 Elements and Noble Gases

Halogens

The halogens are **non-metals**: fluorine, chlorine, bromine, iodine. As you go down the group they become less reactive. It is harder to gain an extra electron because its outer shell is further away from the nucleus. The melting and boiling points also become higher.

Noble Gases

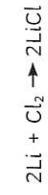
The **noble gases** (**group 0** elements) include: **helium, neon and argon**. They are un-reactive as they have full outer shells, which makes them very stable. They are all colourless gases at room temperature. The boiling points all increase as they go down the group – they have greater intermolecular forces because of the increase in the number of electrons.

Alkali Metals

The alkali metals (**group 1** elements) are soft, very reactive metals. They all have **one electron** in their **outer shell**, making them **very reactive**. They are **low density**. As you go down the group, they become more reactive. They get bigger and it is easier to lose an electron that is further away from the nucleus.

They form ionic compounds with non-metals.

They react with oxygen to form metal oxides.



Metals and Non-metals

They are found at the **left** part of the periodic table. Non-metals are at the **right** of the table.

Metals

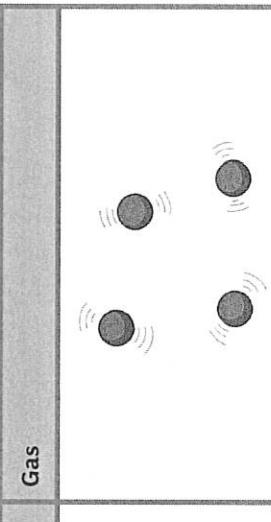
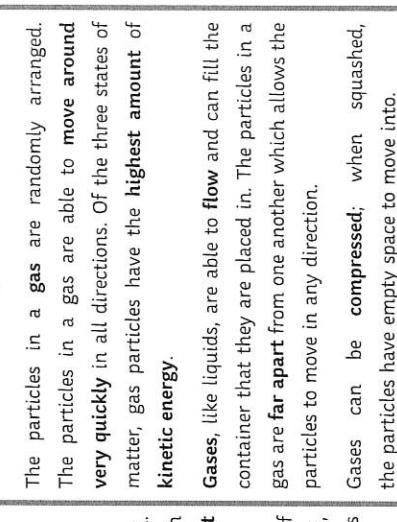
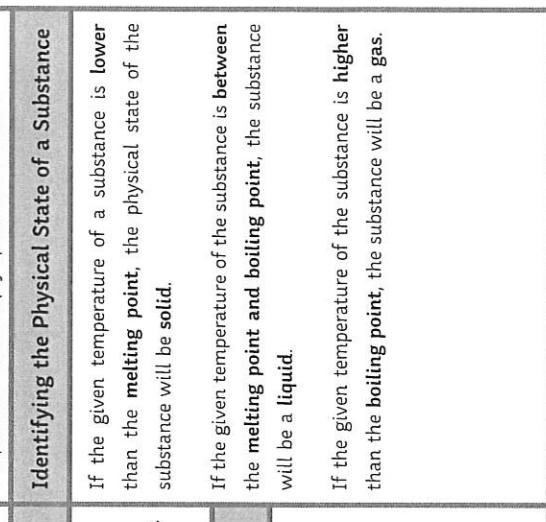
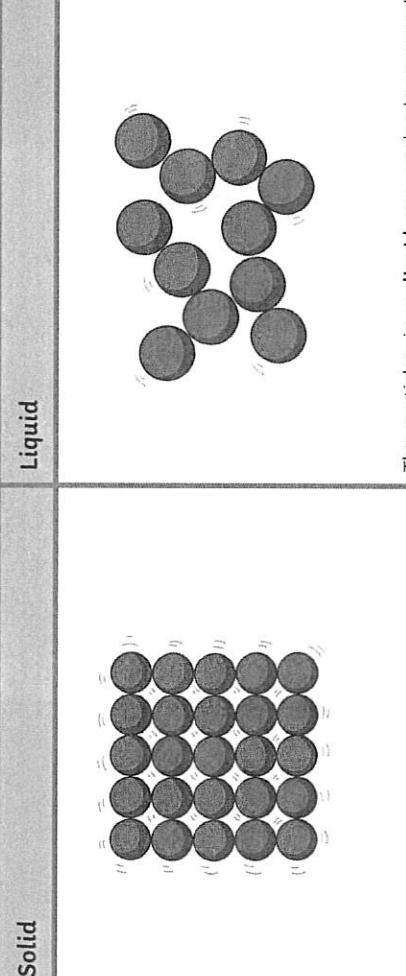
Are strong, malleable, good conductors of electricity and heat. They bond metallically.

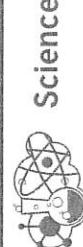
Non-Metals

Are dull, brittle, and not always solids at room temperature.

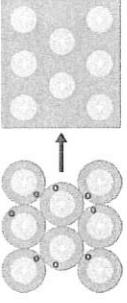
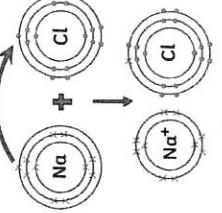
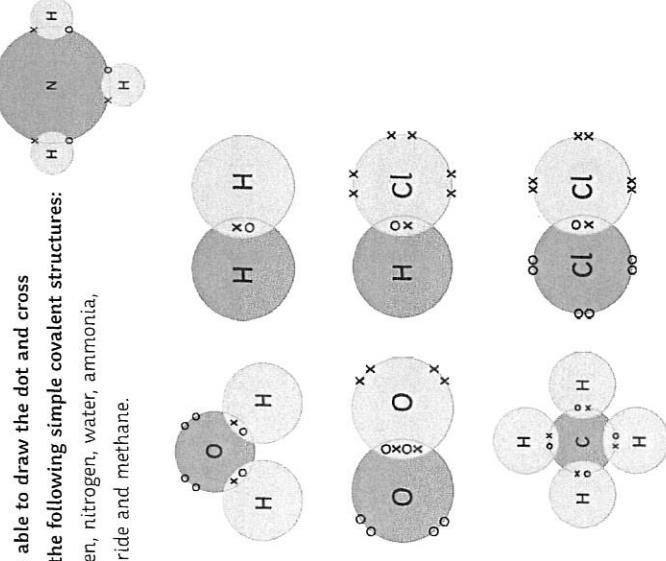
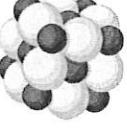


AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter

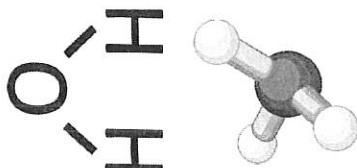
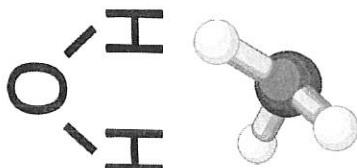
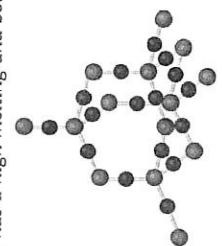
Solid	Liquid	Gas
 <p>The three states of matter are solid, liquid and gas.</p> <p>For a substance to change from one state to another, energy must be transferred.</p> <p>The particles gain energy. This results in the breaking of some of the attractive forces between particles during melting.</p> <p>To evaporate or boil a liquid, more energy is needed to overcome the remaining chemical bonds between the particles.</p>	 <p>The particles in a solid are arranged in a regular pattern. The particles in a solid vibrate in a fixed position and are tightly packed together. The particles in a solid have a low amount of kinetic energy.</p> <p>Solids have a fixed shape and are unable to flow like liquids. The particles cannot be compressed because the particles are very close together.</p> <p>Limitations of the Particle Model (HT only)</p>	 <p>The particles in a gas are randomly arranged. The particles in a gas are able to move around very quickly in all directions. Of the three states of matter, gas particles have the highest amount of kinetic energy.</p> <p>Gases, like liquids, are able to flow and can fill the container that they are placed in. The particles in a gas are far apart from one another which allows the particles to move in any direction.</p> <p>Gases can be compressed; when squashed, the particles have empty space to move into.</p>
 <p>The chemical bonds between particles are not represented in the diagrams above.</p> <p>Particles are represented as solid spheres – this is not the case. Particles like atoms are mostly empty space. Particles are not always spherical in nature.</p>	<p>Identifying the Physical State of a Substance</p> <p>If the given temperature of a substance is lower than the melting point, the physical state of the substance will be solid.</p> <p>If the given temperature of the substance is between the melting point and boiling point, the substance will be a liquid.</p> <p>If the given temperature of the substance is higher than the boiling point, the substance will be a gas.</p>	<p>In chemical equations, the three states of matter are represented as symbols:</p> <ul style="list-style-type: none"> solid (s) liquid (l) gas (g) aqueous (aq) <p>Aqueous solutions are those that are formed when a substance is dissolved in water.</p>
	<p>State Symbols</p>	<p>Science</p>



AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter

Formation of Ions	Metallic Bonding	Ionic Bonding	Covalent Bonding															
<p>Ions are charged particle. They can be either positively or negatively charged, for example Na^+ or Cl^-.</p> <p>When an element loses or gains electrons, it becomes an ion.</p> <p>Metals lose electrons to become positively charged.</p> <p>Non-metals gain electrons to become negatively charged.</p> <p>Group 1 and 2 elements lose electrons and group 6 and 7 elements gain electrons.</p>	<p>Metallic bonding occurs between metals only. Positive metal ions are surrounded by a sea of delocalised electrons. The ions are tightly packed and arranged in rows.</p> 	<p>Ionic bonding occurs between a metal and a non-metal. Metals lose electrons to become positively charged. Opposite charges are attracted by electrostatic forces – an ionic bond.</p> 	<p>Covalent bonding is the sharing of a pair of electrons between atoms to gain a full outer shell. This occurs between non-metals only. Simple covalent bonding occurs between the molecules below. Simple covalent structures have low melting and boiling points – this is because the weak intermolecular forces that hold the molecules together break when a substance is heated, not the strong covalent bonds between atoms. They do not conduct electricity as they do not have any free delocalised electrons.</p> <p>Dot and cross diagrams are useful to show the bonding in simple molecules. The outer electron shell of each atom is represented as a circle, the circles from each atom overlap to show where there is a covalent bond, and the electrons from each atom are either drawn as dots or crosses. There are two different types of dot and cross diagram – one with a circle to represent the outer electron shell and one without.</p> <p>You should be able to draw the dot and cross diagrams for the following simple covalent structures: chlorine, oxygen, water, ammonia, hydrogen chloride and methane.</p> 															
	<table border="1" style="margin-left: auto; margin-right: auto;"> <thead> <tr> <th>Group</th> <th>Ions</th> <th>Element Example</th> </tr> </thead> <tbody> <tr> <td>1</td> <td>+ 1</td> <td>$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$</td> </tr> <tr> <td>2</td> <td>+ 2</td> <td>$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$</td> </tr> <tr> <td>6</td> <td>- 2</td> <td>$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$</td> </tr> <tr> <td>7</td> <td>- 1</td> <td>$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$</td> </tr> </tbody> </table>	Group	Ions	Element Example	1	+ 1	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$	2	+ 2	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$	6	- 2	$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$	7	- 1	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$	<p>Ionic compounds form structures called giant lattices. There are strong electrostatic forces of attraction that act in all directions and act between the oppositely charged ions that make up the giant ionic lattice.</p> 	<p>Properties of Ionic Compounds</p> <ul style="list-style-type: none"> High melting point – lots of energy needed to overcome the electrostatic forces of attraction. High boiling point Cannot conduct electricity in a solid as the ions are not free to move. Ionic compounds, when molten or in solution, can conduct electricity as the ions are free to move and can carry the electrical current.
Group	Ions	Element Example																
1	+ 1	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$																
2	+ 2	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$																
6	- 2	$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$																
7	- 1	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$																
	<p>Metals and Non-metals</p>	<p>Metals are found on the left-hand side of the periodic table. Metals are strong, shiny, malleable and good conductors of heat and electricity. On the other hand, non-metals are brittle, dull, not always solids at room temperature and poor conductors of heat and electricity. Non-metals are found on the right-hand side of the periodic table.</p>																

AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter

Structural Formulae <p>In this type of diagram, the element symbol represents the type of atom and the straight line represents the covalent bonding between each atom.</p>	 Giant Covalent Structure – Graphite <p>Graphite is made up of layers of carbon arranged in hexagons. Each carbon is bonded to three other carbons and has one free delocalised electron that is able to move between the layers. The layers are held together by weak intermolecular forces. The layers of carbon can slide over each other easily as there are no strong covalent bonds between the layers. Graphite has a high melting point because a lot of energy is needed to break the covalent bonds between the carbon atoms. Graphite can conduct electricity.</p>	Nanoscience <table border="1" data-bbox="468 616 642 1099"> <thead> <tr> <th>Name of Particle</th> <th>Diameter</th> </tr> </thead> <tbody> <tr> <td>nanoparticle</td> <td>1–100nm</td> </tr> <tr> <td>fine particles (PM_{2.5})</td> <td>100–2500nm</td> </tr> <tr> <td>coarse particles (PM₁₀)</td> <td>2500–10000nm</td> </tr> </tbody> </table> <p>Fullerenes and Nanotubes</p> <p>Molecules of carbon that are shaped like hollow tubes or balls, arranged in hexagons of five or seven carbon atoms. They can be used to deliver drugs into the body.</p> <p>Buckminsterfullerene has the formula C₆₀</p> <p>Carbon Nanotubes are tiny carbon cylinders that are very long compared to their width. Nanotubes can conduct electricity as well as strengthening materials without adding much weight. The properties of carbon nanotubes make them useful in electronics and nanotechnology.</p> <p>Possible Risks of Nanoparticles</p> <p>As nanoparticles are so small, it makes it possible for them to be inhaled and enter the lungs. Once inside the body, nanoparticles may initiate harmful reactions and toxic substances could bind to them because of their large surface area to volume ratio. Nanoparticles have many applications. These include medicine, cosmetics, sun creams and deodorants. They can also be used as catalysts.</p> <p>Modern nanoparticles are a relatively new phenomenon therefore it is difficult for scientists to truly determine the risks associated with them.</p>	Name of Particle	Diameter	nanoparticle	1–100nm	fine particles (PM _{2.5})	100–2500nm	coarse particles (PM ₁₀)	2500–10000nm
Name of Particle	Diameter									
nanoparticle	1–100nm									
fine particles (PM _{2.5})	100–2500nm									
coarse particles (PM ₁₀)	2500–10000nm									
Structural Formulae <p>In this type of diagram, the element symbol represents the type of atom and the straight line represents the covalent bonding between each atom.</p>	 Giant Covalent Structure – Diamond <p>Each carbon atom is bonded to four other carbon atoms, making diamond very strong. Diamond has a high melting and boiling point. Large amounts of energy are needed to break the strong covalent bonds between each carbon atom. Diamond does not conduct electricity because it has no free electrons.</p>	 <p>Silicon dioxide (silicon and oxygen atoms) has a similar structure to that of diamond, in that its atoms are held together by strong covalent bonds. Large amounts of energy are needed to break the strong covalent bonds therefore silicon dioxide, like diamond, has a high melting and boiling point.</p> 								

AQA GCSE Chemistry (Combined Science) Unit 5.3: Quantitative Chemistry Knowledge Organiser - Foundation

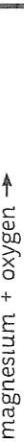
Conservation of Mass

No atoms can be created or made during a chemical reaction, so the mass of the reactants will equal the mass of the product.

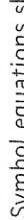
Examples:



Reactions can be shown as a word or symbol equation.



Symbol equations should also be balanced; they should have the same number of atoms on each side.



$$\begin{array}{rcl} \text{H}_2\text{SO}_4 & & \\ \text{A}_r \text{ of H} = 1 & & \text{A}_r \text{ of S} = 32 \\ \text{A}_r \text{ of Cl} = 35.5 & & \text{A}_r \text{ of O} = 16 \\ 1 + 35.5 = 36.5 & & (1 \times 2) + 32 + (16 \times 4) \\ & & 2 + 32 + 64 = 98 \end{array}$$

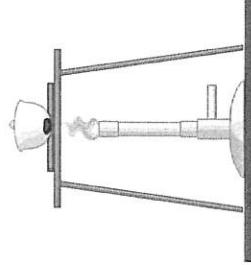
Relative Formula Mass
The relative formulas mass is the sum of all the relative atomic masses of the atoms in the formula.

Calculating Percentage Mass of an Element in a Compound
percentage mass of an element in a compound = $\frac{\text{A}_r \text{ of element}}{\text{M}_r \text{ of the compound}} \times 100\%$

Examples:
 HCl
 $\text{A}_r \text{ of H} = 1$
 $\text{A}_r \text{ of Cl} = 35.5$
 $1 + 35.5 = 36.5$
 H_2SO_4
 $\text{A}_r \text{ of H} = 1$
 $\text{A}_r \text{ of S} = 32$
 $\text{A}_r \text{ of O} = 16$
 $(1 \times 2) + 32 + (16 \times 4)$
 $2 + 32 + 64 = 98$

Calculating Percentage Mass of an Element in a Compound
During a reaction the mass can change. If one of the reactants is a gas, the mass can go up.
E.g.
 $\text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$

Oxygen from the air is added to the magnesium (making the product) which will be heavier in mass.



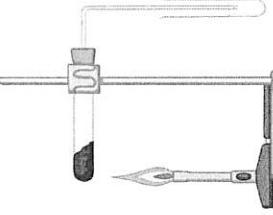
Find the percentage mass of magnesium in magnesium oxide.
 $\text{A}_r \text{ of magnesium} = 24$
 $\text{M}_r \text{ of MgO} = 24 + 16$
 $= 40$
 $\% \text{ mass} = \frac{\text{A}_r}{\text{M}_r} \times 100 = \frac{24}{40} = 0.4 \times 100 = 40\%$

If one of the products is a gas, the mass can go down.
E.g.
 $\text{sodium carbonate} \rightarrow \text{sodium oxide} + \text{carbon dioxide}$

Conservation of Mass
Show that mass is conserved in a reaction.
 $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
 $(2 \times 24) + (2 \times 16) \rightarrow 2(24 + 16)$
 $48 + 32 \rightarrow 2 \times 40$
 $80 \rightarrow 80$

Total M_r on the left-hand side of the equation is the same as the M_r on the right-hand side.

Calculate the mass of the product.
 $8\text{g of magnesium reacts with } 6\text{g of oxygen:}$
 $8 + 6 = 14\text{g of magnesium oxide}$





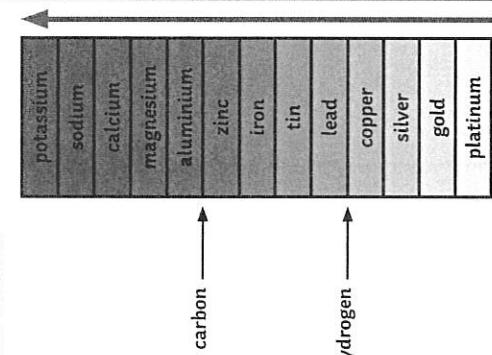
AQA GCSE Chemistry (Combined Science) Unit 4: Chemical Changes Knowledge Organiser

The Reactivity Series

Here's a mnemonic to help you learn the order:

purple (potassium)
slime (sodium)
can (calcium)

make (magnesium)
a (aluminium)
careless (carbon)
zebra (zinc)
insane (iron)
try (tin)
learning (lead)
how (hydrogen)
camels (copper)
surprise (silver)
gorillas (gold)

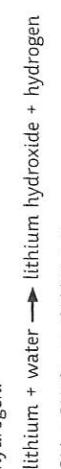


The reactivity series is a league table for metals. The more reactive metals are near the top of the table with the least reactive near the bottom.

In chemical reactions, a more reactive metal will displace a less reactive metal.

Reactions of Metals with Water

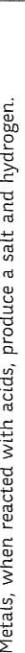
Metals, when reacted with water, produce a metal hydroxide and hydrogen.



The more reactive a metal is, the faster the reaction.

Reactions of Metals with Dilute Acid

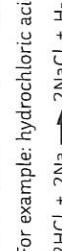
Metals, when reacted with acids, produce a salt and hydrogen.



Metals that are below hydrogen in the reactivity series do not react with dilute acids.

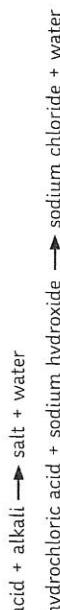
Reactions of Acids

The general formula for the reaction between an acid and a metal is:
acid + metal \rightarrow salt + hydrogen



When an acid reacts with an alkali, a neutralisation reaction takes place and a salt and water are produced.

The general formula for this kind of reaction is as follows:



Naming Salts

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

For example, sodium chloride is formed from the metal carbonate, oxide or hydroxide. The second part of the name comes from the acid that was used to make it. For example, sodium chloride.

Redox Reactions (Higher Tier Only)

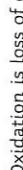
When metals react with acids, they undergo a redox reaction. A **redox reaction** occurs when both **oxidation** and **reduction** take place at the same time.

For example:



The ionic equation can be further split into two half equations.
 $\text{Ca} \rightarrow \text{Ca}^{2+} + 2e^-$

Oxidation is loss of electrons.



Reduction is gaining of electrons.

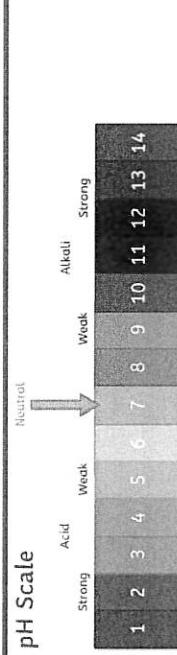
Reactions with Bases

The general formula for the reaction between an acid and a metal oxide is:
acid + metal oxide \rightarrow salt + water



Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:
acid + carbonate \rightarrow salt + water + carbon dioxide



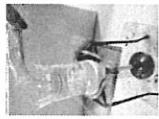
In aqueous solutions, acids produce H⁺ ions and alkalis produce OH⁻ ions.

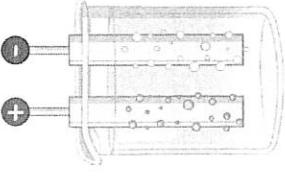
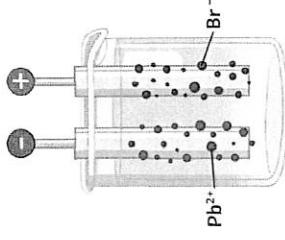
Neutral solutions are pH 7 and are neither acids or alkalis.

For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:
 $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

Making Soluble Salts

1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.
2. Filter the solution to remove the excess copper oxide solid.



AQA GCSE Chemistry (Combined Science) Unit 4: Chemical Changes Knowledge Organiser	
<p>Strong and Weak Acids (Higher Tier Only)</p> <p>A strong acid completely dissociates in a solution.</p> <p>For example: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$</p> <p>Hydrochloric acid is able to completely dissociate in solution to form hydrogen and chloride ions.</p> <p>Examples of strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).</p> <p>Weak acids in comparison only partially dissociate.</p> <p>For example acetic acid partially dissociates to form a hydrogen and acetate ion.</p> $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$ <p>The double arrow symbol indicates that the reaction is reversible. Both the forward and reverse reaction occur at the same time and the reaction never goes to completion.</p>	<p>The Process of Electrolysis</p> <p>Electrolysis is the splitting up of an ionic substance using electricity.</p> <p>On setting up an electrical circuit for electrolysis, two electrodes are required to be placed in the electrolyte. The electrodes are conducting rods. One of the rods is connected to the positive terminal and the other to the negative terminal.</p> <p>The electrodes are inert (this means they do not react in the reaction) and are often made from graphite or platinum.</p> <p>During the process of electrolysis, opposites attract. The positively charged ions will be attracted toward the negative electrode. The negatively charged ions will be attracted towards the positive electrode.</p> <p>When ions reach the electrodes, the charges are lost and they become elements.</p> <p>The positive electrode is called the anode.</p> <p>The negative electrode is called the cathode.</p>
<p>Electrolysis of Aqueous Solutions</p>  <p>Electrolysis of Molten Ionic Compounds – Lead Bromide</p> <p>Lead bromide is an ionic substance. Ionic substances, when solid, are not able to conduct electricity. When molten or in solution, the ions are free to move and are able to carry a charge.</p> <p>The positive lead ions are attracted toward the negative cathode at the same time as the negative bromide ions are attracted toward the positive anode.</p> 	<p>Oxidation is the loss of electrons and reduction is the gaining of electrons. OIL RIG (Higher Tier Only).</p> <p>We represent what is happening at the electrodes by using half equations (Higher Tier Only).</p> <p>The lead ions are attracted towards the negative electrode. When the lead ions (Pb^{2+}) reach the cathode, each ion gains two electrons and becomes a neutral atom. We say that the lead ions have been reduced.</p> $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$ <p>The bromide ions are attracted towards the positive electrode. When the bromide ions (Br^-) reach the anode, each ion loses one electron to become a neutral atom. Two bromine atoms are then able to bond together to form the covalent molecule Br_2.</p> $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$
<p>Using Electrolysis to Extract Metals</p> <p>Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.</p> <p>Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy. This would increase the cost of the process, therefore molten cryolite is added to aluminium oxide to lower the melting point and thus reduce the cost.</p>	

AQA GCSE Chemistry (Combined) Unit 5 Energy Changes Knowledge Organiser

Exothermic and Endothermic Reactions

When a chemical reaction takes place, **energy** is involved. Energy is transferred when chemical bonds are **broken** and when new **bonds are made**.

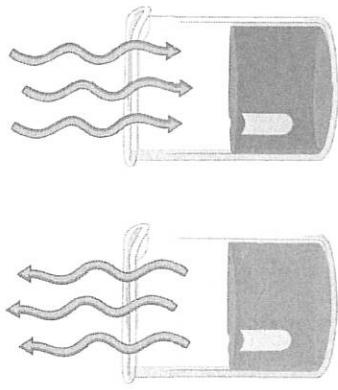
Exothermic reactions are those which involve the transfer of energy **from the reacting chemicals to the surroundings**. During a practical investigation, an exothermic reaction would show an **increase in temperature** as the reaction takes place.

Examples of exothermic reactions include **combustion, respiration and neutralisation** reactions. Hand-warmers and self-heating cans are examples of everyday exothermic reactions.

Endothermic reactions are those which involve the transfer of energy **from the surroundings to the reacting chemicals**. During a practical investigation, an endothermic reaction would show a **decrease in temperature** as the reaction takes place.

Examples of endothermic reactions include the **thermal decomposition** of calcium carbonate.

Eating **sherbet** is an everyday example of an endothermic reaction. When the sherbet dissolves in the saliva in your mouth, it produces a cooling effect. Another example is **instant ice packs** that are used to treat sporting injuries.



Exothermic

Endothermic

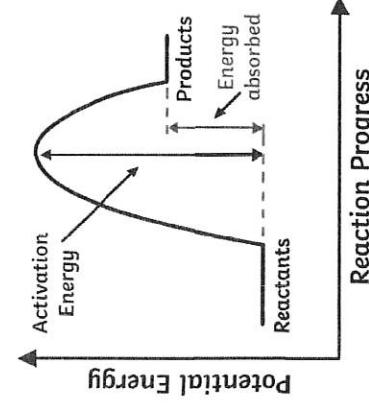
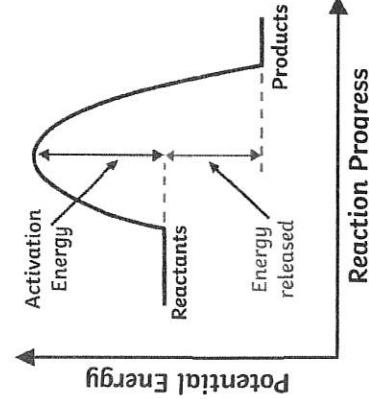
Reaction Profiles – Exothermic

Reaction Profiles – Endothermic

Energy level diagrams show us what is happening in a particular chemical reaction. The diagram shows us the **difference in energy** between the reactants and the products.

In an exothermic reaction, the **reactants** are at a **higher energy level** than the products.

In an exothermic reaction, the difference in energy is **released** to the surroundings and so the **temperature of the surroundings increases**.



Activation Energy – the minimum amount of energy required for a chemical reaction to take place.
Catalysts – increase the rate of a reaction. Catalysts provide an alternative pathway for a chemical reaction to take place by **lowering the activation energy**.

Bond Making and Bond Breaking
In an **endothermic** reaction, energy is needed to break chemical bonds. The **energy change (ΔH)** in an endothermic reaction is **positive**. You may also find, in some textbooks, ΔH referred to as the **enthalpy change**.

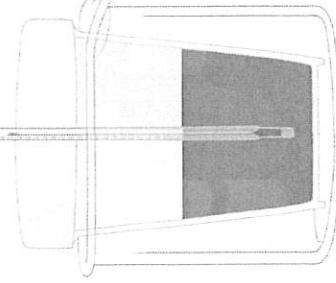
In an **exothermic** reaction, energy is needed to form chemical bonds. The **energy change (ΔH)** in an exothermic reaction is **negative**. Bond energies are measured in **kJ/mol**.

Calculations Using Bond Energies (Higher Tier Only)

Bond energies are used to calculate the change in energy of a chemical reaction.



The first step is to write the symbol equation for the reaction. Once you have done this, work out the bonds that are breaking and the ones that are being made.



On the left-hand side of the equation, the bonds are breaking.

There are two O-H bonds and one O=O bond.

$$\text{So } 464 + 464 + 464 = 1074$$

There are two moles of H_2O_2 therefore the answer needs to be multiplied by two.

$$\text{So } 1074 \times 2 = 2148$$

On the right-hand side of the equation, the bonds are made.

There are two H-O bonds

$$\text{So } 464 + 464 = 928$$

Two moles of H_2O are made therefore the answer needs to be multiplied by two.

$$\text{So } 928 \times 2 = 1856$$

There is also one O=O bond with a bond energy of 498

$$\text{So } 1856 + 498 = 2354$$

$$\Delta H = \text{sum (bonds broken)} - \text{sum (bonds made)}$$

$$\Delta H = 2148 - 2354 = -206 \text{ kJ/mol}$$

The reaction is exothermic as ΔH is negative.

Required Practical

Aim

To investigate the variables that affect temperature changes in reacting solutions, e.g. acid plus metals, acid plus carbonates, neutralisations and displacement of metals.

Equipment

- polystyrene cup
- measuring cylinder
- thermometer
- 250cm^3 glass beaker
- measuring cylinder
- top pan balance

Method

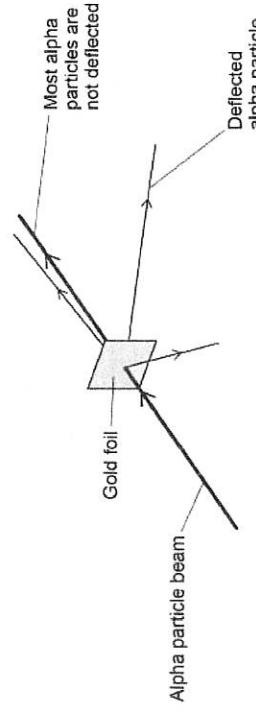
1. Gather the equipment.
2. Place the polystyrene cup inside the beaker. This will prevent the cup from falling over.
3. Using a measuring cylinder, measure out 30cm^3 of the acid. Different acids such as hydrochloric or sulfuric acid may be used. Pour this into the polystyrene cup.
4. Record the temperature of the acid using a thermometer.
5. Using a top pan balance, measure out an appropriate amount of the solid (for example, 10g) or use one strip of a metal such as magnesium.
6. Add the solid to the acid and record the temperature. You may choose to record the temperature of the acid and metal every minute for 10 minutes.

Q1. This question is about gold and compounds of gold.

In the alpha particle scattering experiment alpha particles are fired at gold foil.

Alpha particles are positively charged.

The diagram below shows the results.



(a) Some alpha particles are deflected.

Complete the sentence.

Choose the answer from the box.

negatively charged	not charged	positively charged
--------------------	-------------	--------------------

Some alpha particles are deflected because the nucleus of the atom is _____.

(b) Why are most alpha particles not deflected?

Tick (\checkmark) one box.

The atom is a tiny sphere that cannot be divided.

The atom is mainly empty space.

The electrons orbit the nucleus at specific distances.

(1)

(c) What was one conclusion from the alpha particle scattering experiment?

Tick (\checkmark) one box.

The mass is concentrated at the centre of the atom.

The mass is concentrated at the edge of the atom.

The mass is spread evenly throughout the atom.

(1)

Gold reacts with the elements in Group 7 of the periodic table.

(d) What are Group 7 elements known as?

Tick (\checkmark) one box.

- Alkali metals
Halogens
Noble gases

(1)

(e) Fluorine, chlorine and bromine react with gold.

Which element will be the most reactive with gold?

Tick (\checkmark) one box.

- Fluorine Chlorine Bromine

(f) 3.94 g of gold reacts with chlorine to produce 6.07 g of gold chloride.

The word equation for the reaction is:

gold + chlorine \rightarrow gold chloride

Calculate the mass of chlorine that reacts with 3.94 g of gold.

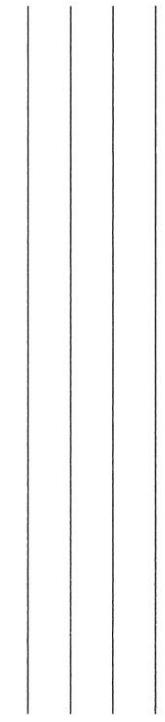
(1)

(g) Calculate the relative formula mass (M_r) of gold chloride (AuCl_3).

Mass = _____ g (1)

Relative atomic masses (A_r): Cl = 35.5 Au = 197

- (b) Figure 2 shows part of the measuring cylinder.



Relative formula mass (M_r) = _____
(2)
(Total 8 marks)

Q2. A student investigated the change in temperature when different masses of zinc were added to copper sulfate solution.

This is the method used.

1. Measure the volume of copper sulfate solution using a measuring cylinder.
2. Pour the copper sulfate solution into a metal container.
3. Add 2 g of zinc.
4. Measure the temperature of the solution.
5. Repeat steps 1 to 4 with different masses of zinc.

Figure 1 shows the apparatus.

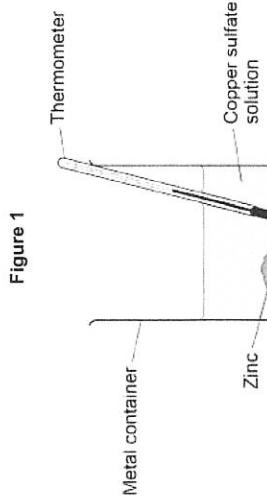
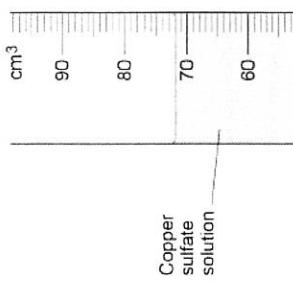


Figure 1

- (a) Give three improvements to the investigation to make the results more accurate.
1 _____
2 _____
3 _____

(3)

Figure 2



Volume = _____ cm³
(1)

What is the volume of copper sulfate solution in Figure 2?

Figure 3

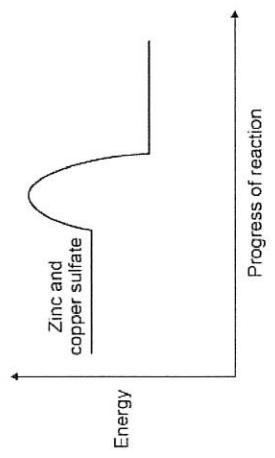


Figure 3

What type of reaction is shown in Figure 3?

Tick (/) one box.

- Endothermic Exothermic Neutralisation
(1)

Figure 4 shows the results.

Figure 4

(2)
(Total 11 marks)

Q3.

This question is about energy changes.

- (a) Which of these items uses an endothermic reaction?

Tick (✓) one box.

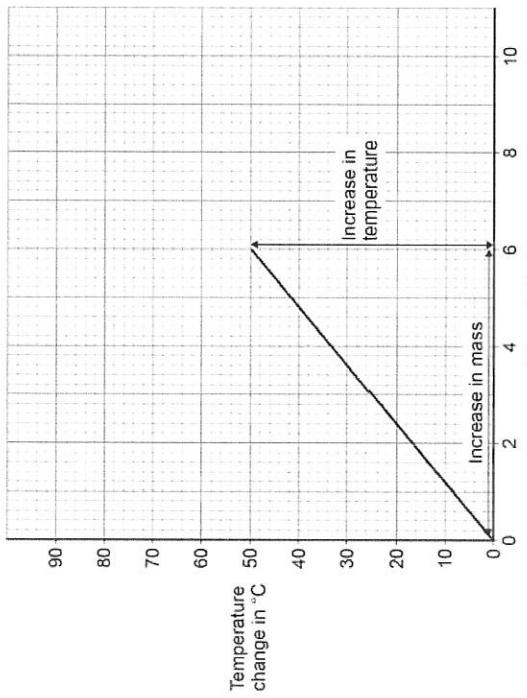
Hand warmer Sports injury pack Self-heating can

Hand warmer

Sports injury pack

Self-heating can

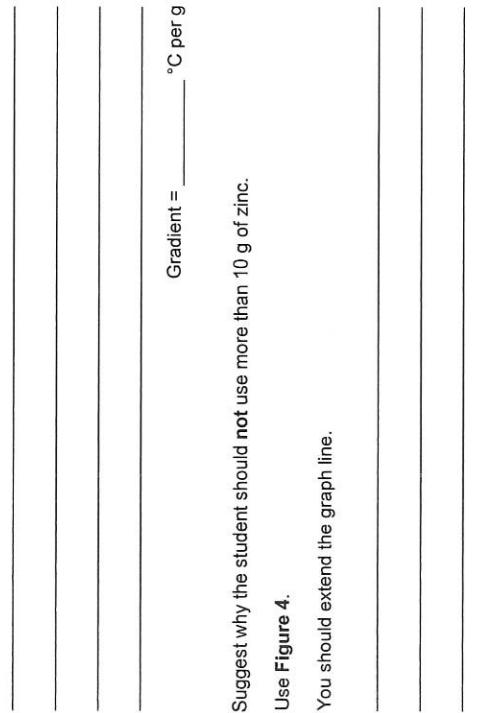
(1)



- (d) Determine the gradient of the line in Figure 4.

Use the equation:

$$\text{gradient} = \frac{\text{increase in temperature in } ^\circ\text{C}}{\text{increase in mass in grams}}$$

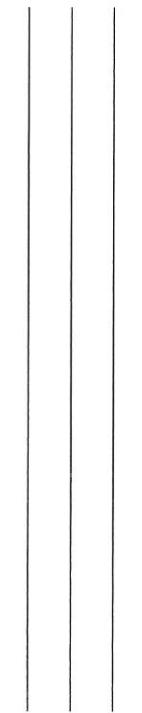


$$\text{Gradient} = \text{_____ } ^\circ\text{C per g}$$

- (e) Suggest why the student should **not** use more than 10 g of zinc.

Use Figure 4.

You should extend the graph line.



- (b) Which letter represents the activation energy for the reaction?

Tick (✓) one box.

A B C D

- (c) Which letter represents the overall energy change for the reaction?

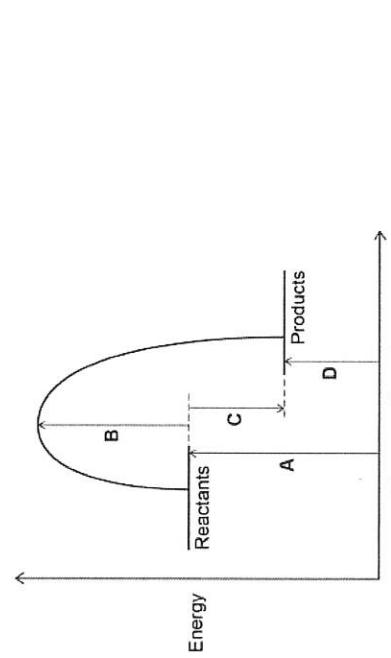


Figure 1 shows the reaction profile for an exothermic reaction.

Figure 1

Tick (✓) one box.

A	<input type="checkbox"/>	B	<input type="checkbox"/>	C	<input type="checkbox"/>	D	<input type="checkbox"/>
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- (d) Complete the sentence.

Choose the answer from the box.

lower than **the same as** **higher than**

In an exothermic reaction the energy of the products

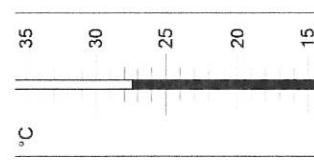
is _____ the energy of the reactants.

- (e) A student measured the temperature at the start and at the end of a reaction.

Name the apparatus used to measure the temperature.

- (f) Figure 2 shows the temperature at the end of the reaction.

Figure 2



Q4.

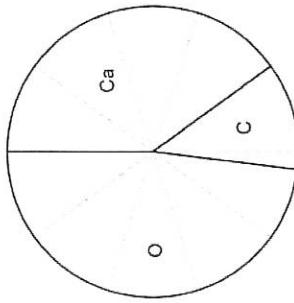
This question is about oxygen and compounds of oxygen.

- (a) What is the state symbol of oxygen at room temperature?

(1)

- (b) Figure 1 shows the percentage by mass of the elements calcium, carbon and oxygen in calcium carbonate.

Figure 1



- (c) What is the percentage by mass of calcium in calcium carbonate?

Percentage = _____ %
(1)

- (d) At high temperature, sodium nitrate decomposes into sodium nitrite and oxygen.

A student heats three samples of sodium nitrate.

The mass of each sample was 4.50 g

The mass of solid after heating was recorded.

- Table 1 shows the mass of solid after heating in each experiment.

Table 1

Experiment	Mass of solid after heating in g
1	3.76
2	3.98
3	4.09

(2)
(Total 7 marks)

Calculate the mean mass of solid after heating.

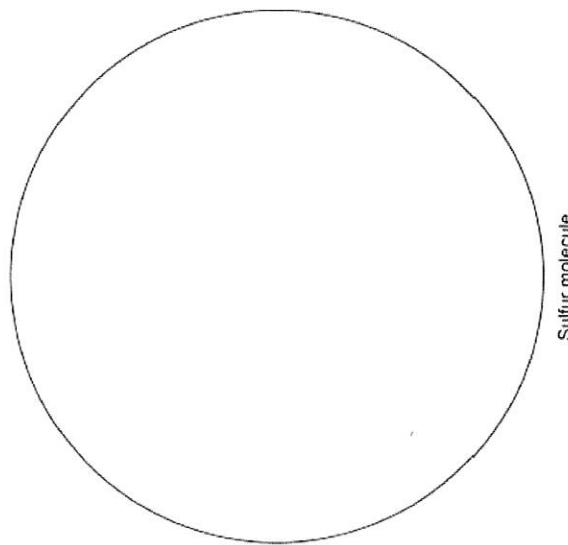
Give your answer to 3 significant figures.

There are _____ bonds between the atoms of oxygen in an oxygen molecule.

(1)

(f) Figure 3 shows the relative sizes of an oxygen molecule and a sulfur molecule.

Figure 3



Mean mass of solid after heating = _____ g

- (d) Table 2 shows the electronic structure of hydrogen and oxygen.

Table 2

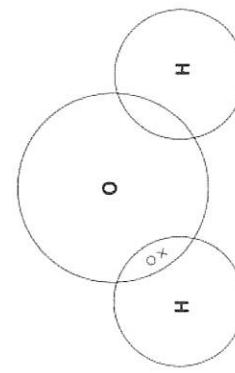
Element	Electronic structure
Hydrogen	1
Oxygen	2,6

Figure 2 shows part of a dot and cross diagram of a molecule of water (H_2O).

Complete the dot and cross diagram.

You should show only the electrons in the outer energy levels.

Figure 2



Oxygen and sulfur are examples of simple molecules.

- (e) Complete the sentence.

Choose the answer from the box.

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

(Total 10 marks)
(2)

- Q5.** This question is about acids and bases.

- (a) What is the pH of sulfuric acid?

Tick (/) one box.

1 7 14

(1)

(b) An acid reacts with zinc to produce zinc chloride and hydrogen.

Which acid reacts with zinc to produce zinc chloride?

Tick () one box.

Hydrochloric acid

Nitric acid

Sulfuric acid

(c) What type of substance is zinc chloride?

Tick () one box.

Alkali

Base

Salt

(d) An alkali is a base in solution.
Which compound is an alkali?

Tick () one box.

Sodium hydroxide

Sodium nitrate

Sodium sulfate

(e) The formula of the copper ion is Cu^{2+}
The formula of the oxide ion is O^{2-}
What is the formula of copper oxide?

Tick () one box.

A student reacts an acid with copper oxide.

(f) The reaction between the acid and copper oxide is very slow at room temperature.

How could the student speed up the reaction?

(g) Complete the sentence to show how the student makes sure that all the acid reacts.

Choose the answer from the box.

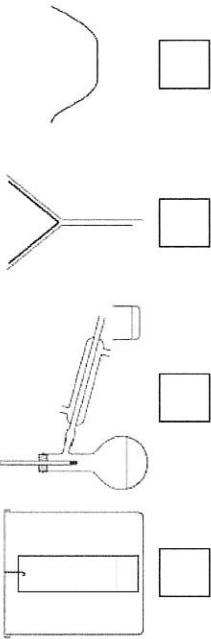
in excess in solution molten soluble

The student adds copper oxide to the acid until the copper oxide is _____.

(h) The student filters the unreacted copper oxide from the solution.

Which apparatus does the student use?

Tick () one box.



(i) What process is used to produce crystals of a salt from a salt solution?

(1)

(Total 9 marks)
(1)

Q6. Magnesium is in Group 2 of the periodic table.

1.0 g of magnesium reacted with chlorine to produce magnesium chloride.

- (a) Which types of element react when magnesium reacted with chlorine?

Tick () one box.

A metal and a metal

A metal and a non-metal

A non-metal and a non-metal

- (b) Write the word equation for the reaction when magnesium reacts with chlorine.



- (c) What apparatus was used to measure the mass of 1.0 g of magnesium?

Tick () one box.

Balance

Beaker

Ruler

- (d) What mass of magnesium chloride was produced?

Tick () one box.

Less than 1.0 g

1.0 g

More than 1.0 g

- (e) Magnesium reacts with oxygen to produce magnesium oxide.

Calculate the percentage mass of magnesium in magnesium oxide (MgO).

Relative atomic mass (A_r): Mg = 24
Relative formula mass (M_r): $\text{MgO} = 40$

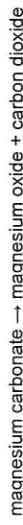
Tick () one box.

A metal and a metal

A metal and a non-metal

Magnesium carbonate decomposes to produce magnesium oxide and carbon dioxide.

The word equation for the reaction is:



Four students heated 2.00 g of magnesium carbonate for 10 minutes.

The table below shows the results.

Mass of carbon dioxide produced in g				
Student 1	Student 2	Student 3	Student 4	Mean
0.97	0.91	0.50	0.95	X

- (f) What is the most likely reason for Student 3's anomalous result?

Tick () one box.

The student heated more than 2.00 g of magnesium carbonate.

The student heated the magnesium carbonate for less than 10 minutes.

The student used a higher temperature.

(g) Calculate value X in the table above.

Do not use the anomalous result.

Give your answer to 2 significant figures.

X (2 significant figures) = _____ 9
 (3)
 (Total 10 marks)

(6)

In graphite the carbon atoms are held together by bonds.

Figure 2 represents part of the structure of graphite.

Q7.

Carbon can exist in a number of different structures.

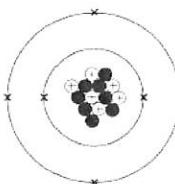
- (a) What is the approximate radius of a carbon atom?

Tick (\checkmark) one box.

0.1 m 0.1 mm 0.1 nm

- (b) Figure 1 shows an atom of carbon.

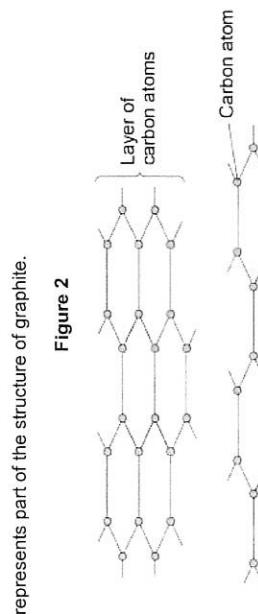
Figure 1



Describe the atomic structure of this carbon atom.

You should include the number of electrons, neutrons and protons.

(6)



- (c) How many bonds does each carbon atom have in graphite?

Use Figure 2.

Tick (\checkmark) one box.

1	<input type="checkbox"/>	2	<input type="checkbox"/>	3	<input type="checkbox"/>	4	<input type="checkbox"/>
---	--------------------------	---	--------------------------	---	--------------------------	---	--------------------------

(1)

- (d) What type of bonds hold the carbon atoms together in graphite?

Tick (\checkmark) one box.

Covalent	<input type="checkbox"/>
Ionic	<input type="checkbox"/>
Metallic	<input type="checkbox"/>

(1)

- (e) Lubricants allow objects to slide over each other easily.

Suggest why graphite can be used as a lubricant.

Use Figure 2.

(1)

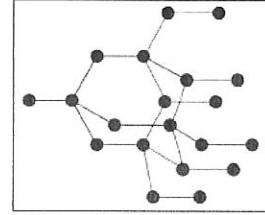
(f) The two structures represent different forms of carbon.

Draw one line from each structure to the form of carbon.

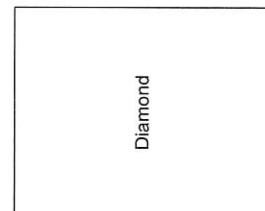
Structure

Form of carbon

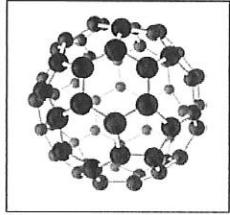
Buckminsterfullerene



Diamond



Graphene



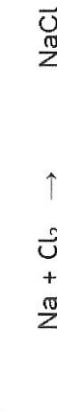
Nanotube

(Total 12 marks)

Q8. This question is about Group 1 elements.

(a) Sodium reacts with chlorine to produce sodium chloride.

Balance the equation for the reaction.



(1)

(b) 4.6 g of sodium reacts with chlorine to produce 11.7 g of sodium chloride.

What mass of chlorine reacted?

Mass of chlorine = _____ g (1)

(c) A teacher puts hot sodium into a gas jar of chlorine.

The changes seen before, during and after this reaction were observed.

Complete the sentences.

Choose the answers from the box.

colourless	green	liiac	silver	white
yellow				

Sodium is a _____ solid.

Chlorine is a _____ gas.

The hot sodium burns with a _____ flame.

The product sodium chloride is a _____ solid.

(d) Sodium chloride (NaCl) is an ionic compound.

Write the formulae of the ions in sodium chloride.

Sodium ion _____

Chloride ion _____

(e) Complete the sentence.

Choose the answer from the box.

an atom	an electron	a neutron	a proton
---------	-------------	-----------	----------

Potassium is more reactive than sodium.

This is because potassium loses _____ more easily than sodium.

(f) How does the size of a potassium atom compare with the size of a sodium atom?

Give a reason for your answer.

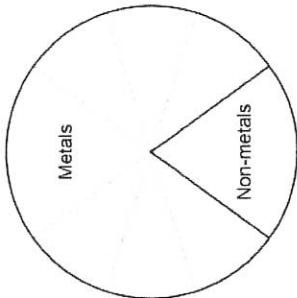
(1)

Reason _____

Q9. (Total 11 marks)

This question is about elements and compounds.

- (a) The chart below shows the proportion of elements in the periodic table that are metals and non-metals.



Determine the percentage of the elements in the chart above that are metals.

$$\text{Percentage} = \frac{\text{Area of Metal Sector}}{\text{Total Area}} \times 100\%$$

- (b) Give **two** physical properties of metals.

1. _____
2. _____

(2)

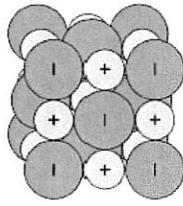
- (c) Sodium reacts with chlorine to produce sodium chloride.

Balance the equation for the reaction.



(1)

The diagram below shows part of the structure of sodium chloride (NaCl).

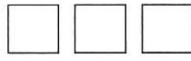


Sodium chloride

- (d) What holds the particles together in sodium chloride?

Use the diagram above.

Tick () **one** box.



Electrostatic attractions

Intermolecular forces

Metallic bonds

(1)

- (e) Solid sodium chloride does not conduct electricity.

Give **two** ways in which sodium chloride can be made to conduct electricity.

1. _____
2. _____

(2)
(Total 8 marks)

Q10.

This question is about electrolysis.

- (a) Complete the sentence.

Choose the answer from the box.

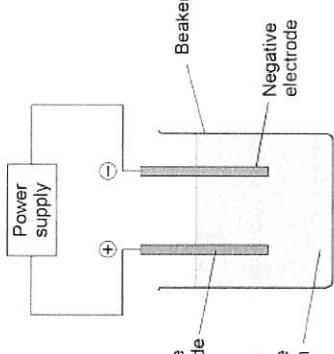
<input type="checkbox"/> gaseous	<input type="checkbox"/> molten	<input type="checkbox"/> solid
----------------------------------	---------------------------------	--------------------------------

Copper chloride can conduct electricity when in solution or when _____.

(1)



Figure 1



There are four ions in copper chloride solution:

- * Cu^{2+}
- * Cl^-
- * H^+
- * OH^-

(b) Why do Cl^- ions and OH^- ions move to the positive electrode?

(1)

(c) Where do the H^+ and OH^- ions come from in the electrolysis of copper chloride solution?

Tick (✓) one box.

Air

Copper chloride

Water

(d) Which ion produces a metal?

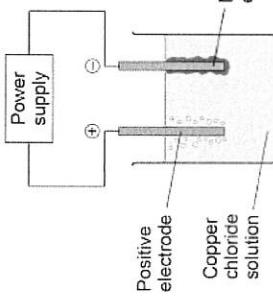
Tick (✓) one box.



(1)

(e) Figure 1 shows the apparatus used for the electrolysis of copper chloride solution.

Figure 2



(e) Figure 2 shows the apparatus during the electrolysis of copper chloride solution.

Describe what is seen at each electrode during the electrolysis of copper chloride solution.

Positive electrode _____

Negative electrode _____

(f) 500 cm³ of copper chloride solution contains 6.50 g of copper chloride.

Calculate the mass of copper chloride in 40.0 cm³ of this copper chloride solution.

(1)

Mass = _____ g
(Total 8 marks)

Q11.

This question is about elements in the periodic table.

- (a) What property was used to arrange elements in early periodic tables?

Tick (\checkmark) one box.

Gas	<input type="checkbox"/>
Liquid	<input type="checkbox"/>
Solid	<input type="checkbox"/>

(1)

Atomic number

<input type="checkbox"/>
<input type="checkbox"/>
<input type="checkbox"/>

Atomic weight

Atomic weight

<input type="checkbox"/>
<input type="checkbox"/>
<input type="checkbox"/>

Mass number

(1)

- (b) In early periodic tables, iodine (I) was placed before tellurium (Te).

Mendeleev placed iodine after tellurium.

Figure 1 shows part of Mendeleev's periodic table.

Figure 1

16	19
0	F
32	35.5
S	Cl
79	80
Se	Br
128	127
Te	I

Suggest one reason why Mendeleev placed iodine in the column shown in Figure 1.

The table below shows the melting points of three Group 1 metals.

Metal	Melting point in °C
Lithium	180
Sodium	98
Potassium	63

- (c) What state is lithium at 100 °C?

Use table above.

Tick (\checkmark) one box.

Gas	<input type="checkbox"/>
Liquid	<input type="checkbox"/>
Solid	<input type="checkbox"/>

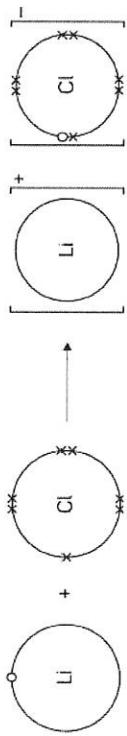
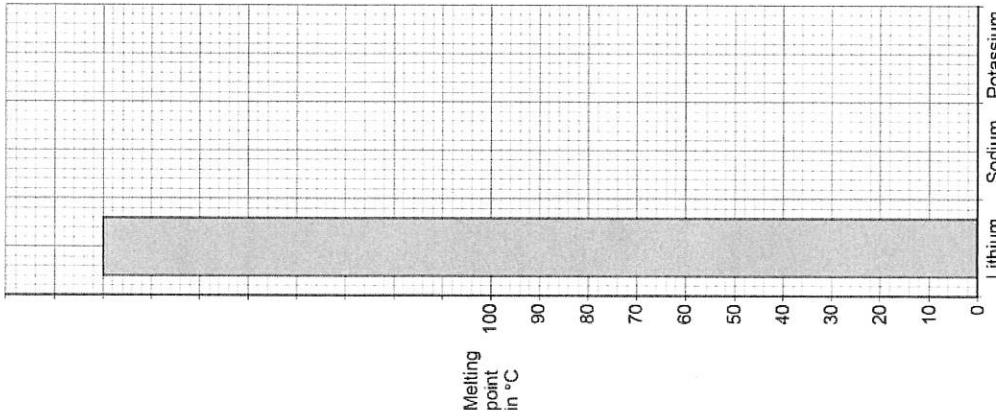
Complete the graph below.

Use the table above.

You should:

- complete the scale on the y-axis
- draw bars to show the melting points of sodium and potassium.

(1)



Describe what happens to a lithium atom and to a chlorine atom when they react.

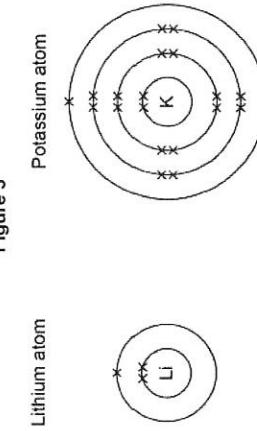
Use Figure 2 to answer in terms of electrons.

(3) _____

(f) Lithium and potassium are in the same group of the periodic table.

Figure 3 represents the electronic structures of a lithium atom and of a potassium atom.

Figure 3



Give two reasons why potassium is more reactive than lithium.

1. _____
 2. _____
-
-

(2)
(Total 11 marks)

Q12.

A student investigated the temperature change when metal X was added to copper sulfate solution.

This is the method used.

1. Add 25 cm³ of copper sulfate solution to a beaker.
2. Measure the temperature of the copper sulfate solution.
3. Add 1.0 g of metal X and stir.
4. Measure the highest temperature reached when metal X is added to copper sulfate solution.
5. Repeat steps 1 to 4 with different metals.

Figure 1 shows the apparatus used.

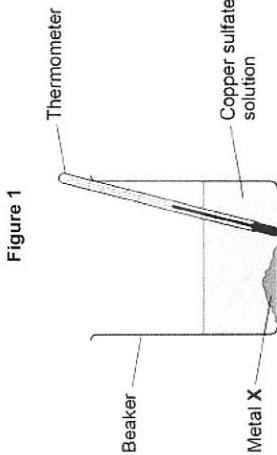
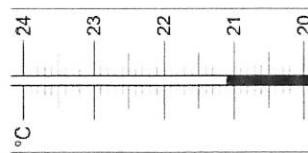


Figure 1

Figure 2 shows the thermometer reading of the copper sulfate solution at the start of the investigation.

Figure 2



Temperature change in °C	Test 1	Test 2	Test 3	Test 4
	9.2	7.3	9.5	9.2

Calculate the mean temperature change for metal Y.

Do not include the anomalous result in your calculation.

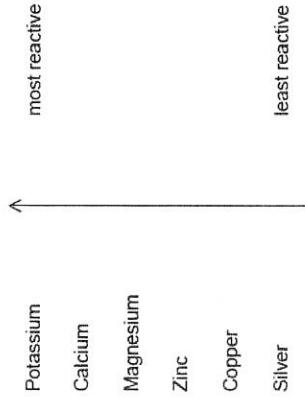
Mean temperature change =	_____ °C
---------------------------	----------

The more reactive the metal added to copper sulfate solution, the greater the temperature change.

Figure 3 shows a reactivity series.

Figure 3

- (a) The highest temperature reached when metal X was added to copper sulfate solution was 35.5 °C
- Determine the temperature change when metal X is added to copper sulfate solution.
- Use Figure 2.



(d) The student repeated the experiment.

The student added:

- magnesium to copper sulfate solution
- an unknown metal A to copper sulfate solution.

Table 2 shows the results.

Table 2

Metal	Temperature change in °C
Magnesium	12
Metal A	8

The student concludes metal A is zinc.

Give one reason why the student is correct.

Use Figure 3 and Table 2.

Increases	<input type="checkbox"/>
Stays the same	<input type="checkbox"/>
Decreases	<input type="checkbox"/>

(e) The student did the experiment with silver and copper sulfate solution.

What happens to the temperature of the mixture?

Use Figure 3.

Tick (/) one box.

(f) Suggest one reason why the student should not add potassium metal to copper sulfate solution.

<input type="checkbox"/>
<input type="checkbox"/>
<input type="checkbox"/>

(g) 100 cm³ of the copper sulfate solution contains 1.8 g of copper sulfate.

Calculate the mass of copper sulfate in 25 cm³ of this copper sulfate solution.

Mass = _____ g
(Total 11 marks)

Q13.

Sodium and potassium are Group 1 elements.

(a) What is the name of Group 1 elements?

Tick (/) one box.

<input type="checkbox"/>
<input type="checkbox"/>
<input type="checkbox"/>

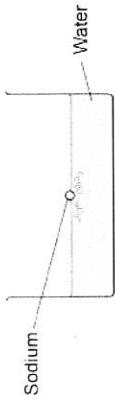
Alkali metals

Halogens

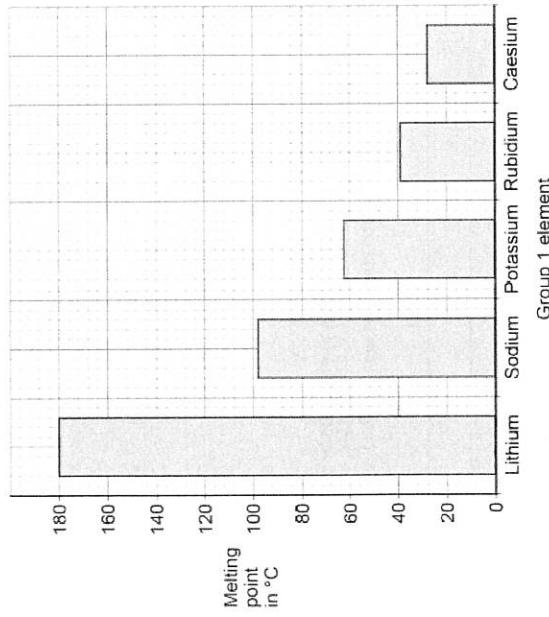
Noble gases

(b) Figure 1 represents the melting points of Group 1 elements.

Figure 1



Compare what is seen when sodium reacts with water and when potassium reacts with water.

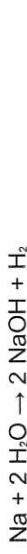


What is the melting point of sodium?

Melting point of sodium = _____ °C

- (c) Sodium reacts with water to produce sodium hydroxide and hydrogen.

Balance the equation for the reaction.



- (d) Calculate the relative formula mass (M_r) of sodium hydroxide (NaOH).

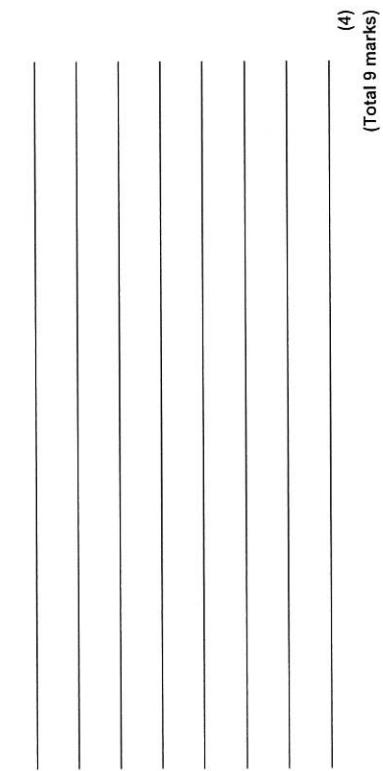
Relative atomic masses (A_r): H = 1 O = 16 Na = 23

Relative formula mass (M_r) = _____
(2)

- (e) Sodium and potassium both react with water.

Figure 2 shows sodium reacting with water.

Figure 2



(4)
(Total 9 marks)