Chemical Changes and Energy Changes Revision Booklet

Chemistry Paper 1

AQA TRILOGY Chemistry (8464) from 2016 Topics T5.4 Chemical changes				
Topic	**Student Checklist**	**R**	**A**	**G**
5.4.1 Reactivity of metals	Describe how metals react with oxygen and state the compound they form, define oxidation and reduction			
	Describe the arrangement of metals in the reactivity series, including carbon and hydrogen, and use the reactivity series to predict the outcome of displacement reactions			
	Recall and describe the reactions, if any, of potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper with water or dilute acids			
	Relate the reactivity of metals to its tendency to form positive ions and be able to deduce an order of reactivity of metals based on experimental results			
	Recall what native metals are and explain how metals can be extracted from the compounds in which they are found in nature by reduction with carbon			
	Evaluate specific metal extraction processes when given appropriate information and identify which species are oxidised or reduced			
5.4.2 Reactions of acids	**HT ONLY: Describe oxidation and reduction in terms of loss and gain of electrons**			
	HT ONLY: Write ionic equations for displacement reactions, and identify which species are oxidised and reduced from a symbol or half equation			
	HT ONLY: Explain in terms of gain or loss of electrons that the reactions between acids and some metals are redox reactions, and identify which species are oxidised and which are reduced (Mg, Zn, Fe + HCl & H_2SO_4)			
	Explain that acids can be neutralised by alkalis, bases and metal carbonates and list the products of each of these reactions			
	Predict the salt produced in a neutralisation reaction based on the acid used and the positive ions in the base, alkali or carbonate and use the formulae of common ions to deduce the formulae of the salt			
	Describe how soluble salts can be made from acids and how pure, dry samples of salts can be obtained			
	Required practical 8: *preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution*			
	Recall what the pH scale measures and describe the scale used to identify acidic, neutral or alkaline solutions			
	Define the terms acid and alkali in terms of production of hydrogen ions or hydroxide ions (in solution), define the term base			
	Describe the use of universal indicator to measure the approximate pH of a solution and use the pH scale to identify acidic or alkaline solutions			
	HT ONLY: Use and explain the terms dilute and concentrated (in terms of amount of substance) and weak and strong (in terms of the degree of ionisation) in relation to acids			
	HT ONLY: Explain how the concentration of an aqueous solution and the strength of an acid affects the pH of the solution and how pH is related to the hydrogen ion concentration of a solution			
5.4.3 Electrolysis	Describe how ionic compounds can conduct electricity when dissolved in water and describe these solutions as electrolytes			
	Describe the process of electrolysis			
	Describe the electrolysis of molten ionic compounds and predict the products at each electrode of the electrolysis of binary ionic compounds			
	Explain how metals are extracted from molten compounds using electrolysis and use the reactivity series to explain why some metals are extracted with electrolysis instead of carbon			
	Describe the electrolysis of aqueous solutions and predict the products of the electrolysis of aqueous solutions containing single ionic compounds			
	Required practical 9: *investigate what happens when aqueous solutions are electrolysed using inert electrodes*			

		R	A	G
	HT ONLY: Describe the reactions at the electrodes during electrolysis as oxidation and reduction reactions and write balanced half equations for these reactions			

AQA TRILOGY Chemistry (8464) from 2016 Topics T5.5 Energy changes				
Topic	**Student Checklist**	**R**	**A**	**G**
5.5.1 Exothermic and endothermic reactions	Describe how energy is transferred to or from the surroundings during a chemical reaction			
	Explain exothermic and endothermic reactions on the basis of the temperature change of the surroundings and give examples of everyday uses			
	Required practical 10*: investigate the variables that affect temperature changes in reacting solutions*			
	Describe what the collision theory is and define the term activation energy			
	Interpret and draw reaction profiles of exothermic and endothermic reactions, inc identifying the relative energies of reactants and products, activation energy and overall energy change			
	HT ONLY: Explain the energy changes in breaking and making bonds and calculate the overall energy change using bond energies			

5.4.1 Reactivity of Metals

A **metal compound** within a **rock** is an **ore.** Ores are **mined** and then **purified**.

Whether it is worth extracting a particular metal depends on:

- **How easy it is to extract it from its ore**
- **How much metal the ore contains**
- **The changing demands for a particular metal**

Most **metals** in ores are **chemically bonded** to **other elements** in compounds. Many of these metals have been **oxidised** (have oxygen added) by oxygen in the air to form their oxides.

Iron + oxygen → iron (III) oxide

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

To extract metals from their oxides, the metal oxides must be **reduced** (have oxygen removed).

Metals can be arranged in order of reactivity in a **reactivity series**.

Order of reactivity	Reaction with water	Reaction with acid
Potassium	Fizz, giving off hydrogen gas and leaving an alkaline solution of metal hydroxide	Reacts violently and explodes
Sodium		
Lithium		
Calcium		Fizz, giving off hydrogen gas and forming a salt
Magnesium	Very slow reaction	
Aluminium		
Zinc		
Iron		
Tin	No reaction with water at room temperature	React slowly with warm acid
Lead		
Copper	No reaction	No reaction
Silver		
Gold		

Metals can be arranged in order of reactivity in a **reactivity series**.

When metals react with other substances the metal atoms form **positive ions**.

The reactivity of a metal is linked to its tendency to form positive ions.

The **non-metals hydrogen** and **carbon** are often included in the series as they can be used to extract less reactive metals.

Metal + acid → salt + hydrogen

Potassium
Sodium
Lithium
Calcium
Magnesium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold

The **reactivity** of a metal determines the **method** of **extraction.**

Metals **above** carbon must be extracted from their ores by using **electrolysis.**

Metals **below** carbon can be extracted from their ores by **reduction** using **carbon.**
REDUCTION involves the loss of oxygen.

metal oxide + carbon → metal + carbon dioxide

Gold and **silver** do not need to be extracted. They occur **native** (naturally).

Potassium
Sodium
Calcium
Magnesium
Aluminium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold
Platinum

Increasing reactivity

Oxidation and reduction in terms of electrons (HT)

A **more reactive metal can displace a less reactive metal from its compound** in **displacement reactions.**

Iron + copper(II) sulfate → iron sulfate + copper

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

Higher:

OILRIG

Oxidation **I**s **L**oss of electrons
Reduction **I**s **G**ain of electrons

When reactions involve **oxidation and reduction**, they are known as **redox reactions**

Higher:

An **ionic equation** shows only the atoms and ions that change in a reaction:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

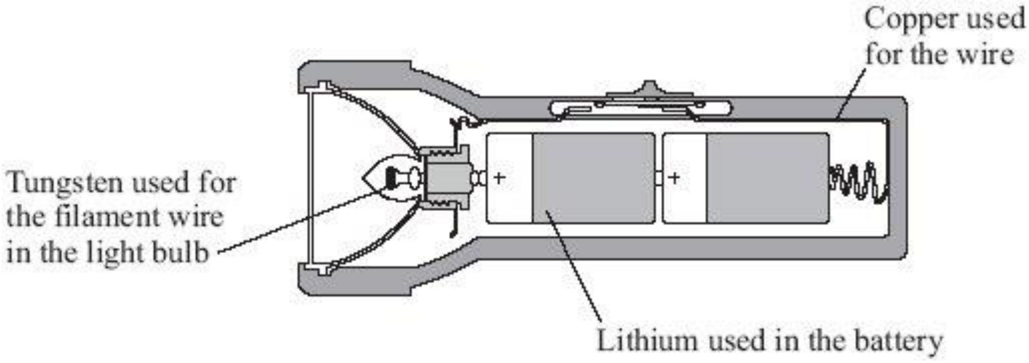
Half equations show what happens to each reactant:

$Fe \rightarrow Fe^{2+} + 2e^-$ — *The iron atoms are **oxidised** (lose 2 electrons) to form **ions**.*

$Cu^{2+} + 2e^- \rightarrow Cu$ — *The 2 **electrons** from the iron are **gained** (**reduction**) by copper ions as they become **atoms**.*

Q1.

The diagram shows a circuit that is used in a torch. Electrons flow through this circuit.

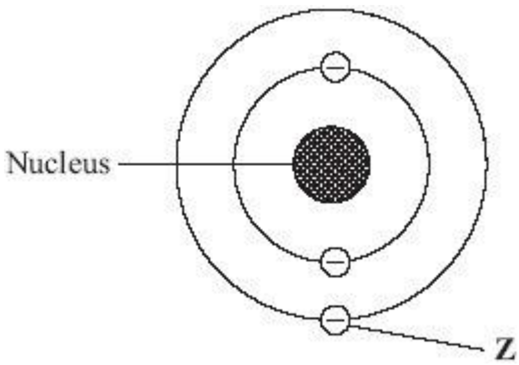


(a) Why is copper used for the wire?

____

(1)

(b) The diagram shows the structure of an atom of lithium.



Name the particle labelled **Z**.

____

(1)

(c) The table shows some properties of the metals used in the electrical circuit.

Metal	Melting point in °C	Boiling point in °C	Reaction with oxygen
Copper	1083	2582	Reacts **slowly** to form a thin oxide layer on surface
Lithium	179	1317	Reacts **rapidly** to form oxide
Tungsten	3370	5930	Reacts **only** when very

			hot to form oxide

(i) Use information from the table to suggest the order of reactivity for copper, lithium and tungsten.

most reactive ________________________________

____

least reactive ________________________________

(2)

(ii) The filament wire glows because it gets very hot.

Use information from the table to suggest **one** reason why tungsten is used for the filament wire in the light bulb.

____

____

(1)

(d) The gas used in the light bulb is argon.

Draw a ring around the correct word in the box to complete the sentence.

Argon is used in the light bulb because it is	dense. solid. unreactive.	.

(1)

(Total 6 marks)

Q2.

The table gives information about some metals.

Name of the metal	Cost of one tonne of the metal in December 2003 (£)	Percentage of the metal in the crust of the earth (%)
Aluminium	883	8.2
Platinum	16720000	0.0000001
Iron	216	4.1
Gold	8236800	0.0000001

(a) Use information in the table to suggest why gold and platinum are very expensive metals.

____

____

(1)

(b) Aluminium and iron are made by *reduction* of their ores.

(i) Name the element that is removed from the ores when they are *reduced.*

____

(1)

(ii) Use the reactivity series on the Data Sheet to suggest a metal that would reduce aluminium ore.

____

(1)

(c) Aluminium is made by the reduction of molten aluminium ore, using a very large amount of electricity.

(i) How is iron ore reduced in a blast furnace to make iron?

____

____

____

____

(2)

(ii) Suggest why aluminium is more expensive than iron.

____

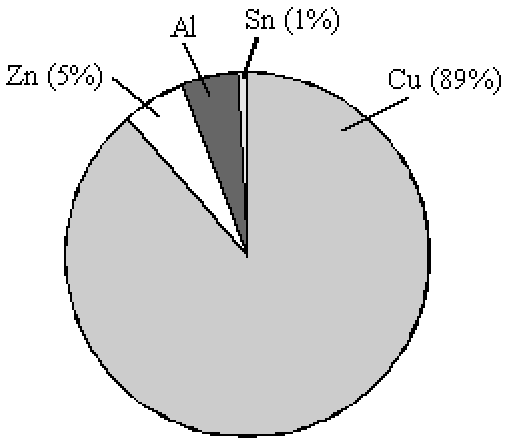
____

(1)
(Total 6 marks)

Q3.

The 50 Eurocent coin is made from an alloy called 'Nordic Gold'.

The pie chart shows the percentage by mass of each metal in 'Nordic Gold'.



(a) (i) Calculate the percentage of aluminium, Al, in the coin.

____

(1)

(ii) The 50 Eurocent coin has a mass of 7 grams.
Calculate the mass of zinc, Zn, in this coin.

____

____

Mass of zinc = ______________________ g

(2)

(b) Zinc is extracted by removing oxygen from zinc oxide.

(i) What name is given to a reaction in which oxygen is removed from a substance?

____

(1)

(ii) Explain how oxygen can be removed from zinc oxide to make zinc. Use the reactivity series on the Data Sheet to help you

____

____

____

____

(2)

(Total 6 marks)

5.4.2 Reactions of Acids

Acids react with some **metals** to produce **salts** and **hydrogen**.

Metal + acid → salt + hydrogen

Reactions between metals and acids only occur if the metal is **more reactive** than the **hydrogen** in the acid. If the metal is too reactive, the reaction with acid is **violent**.

The **salt** that is made depends on the **metal** and **acid** used.

Salts made when metals react *nitric acid* are called *nitrates*.

Zinc + Nitric acid → Zinc Nitrate + Hydrogen

Salts made when metals react with *sulfuric acids* are called *sulfates*.

Iron + Sulfuric Acid → Iron Sulfate + Hydrogen

Salts made when metals react with *hydrochloric* acid are called *chlorides*.

Magnesium + Hydrochloric acid → Magnesium Chloride + Hydrogen

(HT only)

In the reaction between magnesium and hydrochloric acid, the hydrogen ions are **displaced** from the solution by magnesium as the magnesium is more reactive than hydrogen.

The following **ionic equation** occurs:

$$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

The chloride ions are not included as they do not change in the reaction. These are known as **spectator ions.**

The reaction can be further represented by **half equations,** showing that the reaction between a metal and acid is a **redox reaction.**

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

The magnesium atoms lose two electrons, they have been **oxidised**.

$$2H^{+} + 2e^{-} \rightarrow H_2$$

The hydrogen ions have gained electrons, they have been **reduced**.

Acids are neutralised by **alkalis** (eg: soluble metal hydroxides) and **bases** (eg: insoluble metal hydroxides and metal oxides) to produce **salts** and **water** and by **metal carbonates** to produce **salts, water** and **carbon dioxide**.

The salt **name** depends on the **acid** used and the **positive ions** in the **alkali, base or carbonate**.

Making Soluble Salts from acids and alkalis

Salts can be made by reacting an acid with an alkali.

Acid + Alkali → Salt + Water

Making Soluble Salts from acids and bases

Salts can be made by reacting an acid with a insoluble base.

Acid + Bases → Salt + Water

Making Soluble Salts from acids and metal carbonates

Salts can be made by reacting an acid with a metal carbonate.

Acid + Metal carbonate → Salt + Water + Carbon dioxide

Salts are made of positive metal ions (or ammonia ions - NH_4^+) and a negative ion from the acid. Like all ionic compounds, salts have **no overall charge**, so once you know the charges on the ions, you can work out the **formula**.
Example: **magnesium sulfate** is **$MgSO_4$**

ion	formula	ion	formula
Group 1	Li^+ Na^+ K^+	Transition metals	Cu^{2+} Fe^{3+}
Group 2	Mg^{2+} Ca^{2+}	Group 7	F^- Cl^- Br^-
Aluminium	Al^{3+}	Nitrate	NO_3^-
Ammonium	NH_4^+	Sulphate	SO_4^{2-}

Required Practical

Soluble salts can be made from **acids** by reacting them with **solid insoluble substances**, such as **metals, metal oxides, hydroxides or carbonates**. The solid is added to the acid until no more reacts and the excess solid is **filtered** off to produce a solution of the salt. Salt solutions can be **crystallised** to produce **solid** salts. **You will complete this as a required practical.**

1. Measure the required volume of acid with a measuring cylinder and add the weighed solid (insoluble metal, oxide, hydroxide or carbonate) in small portions with stirring.
2. Safety goggles required - the mixture may be heated to speed up the reaction. When no more of the solid dissolves it means ALL the acid is neutralised and there should be a little excess solid. You should see a residue of the solid (oxide, hydroxide, carbonate) left at the bottom of the beaker.
3. Filter the solution to remove the excess solid metal/oxide/carbonate, into an evaporating dish. On filtration, only a solution of the salt is left.
4. Then hot concentrated solution is left to cool and crystallise.
   After **crystallisation**, you collect and dry the crystals with a filter paper.
   If the solution is heated, the solvent will evaporate faster. Heating a solution until all the solvent has evaporated is known as **heating to dryness**.

Indicators are substances which change **colour** when you add them to acids and alkali.

Litmus goes red in acid and blue in alkali.

Universal indicator, made from many dyes is used to tell you **pH**. The scale runs from 0 (most acidic) to 14 (most alkaline). Aqueous solutions of **acids** have a pH value **less than 7**, and for **alkalis greater than 7** and anything in the middle is **neutral** (pH 7). You can use a pH meter to record the change of a pH over time.

Acids produce **hydrogen ions (H^+)** in aqueous solutions and **alkalis** produce **hydroxide ions (OH^-).** In **neutralisation** reactions between an acid and alkali, hydrogen ions react with hydroxide ions to produce **water.**

Neutralisation symbol equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

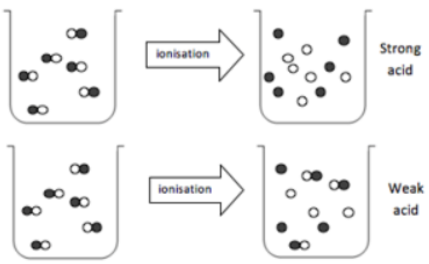
pH	Examples of solutions
0	Battery acid, strong hydrofluoric acid
1	Hydrochloric acid secreted by stomach lining
2	Lemon juice, gastric acid, vinegar
3	Grapefruit juice, orange juice, soda
4	Tomato juice, acid rain
5	Soft drinking water, black coffee
6	Urine, saliva
7	"Pure" water
8	Sea water
9	Baking soda
10	Great Salt Lake, milk of magnesia
11	Ammonia solution
12	Soapy water
13	Bleach, oven cleaner
14	Liquid drain cleaner

Acids must **dissolve** in water to show their acidic properties.
A **concentrated acid** has a relatively **large amount of solute** dissolved in the solvent.
A **dilute acid** has a relatively **smaller amount of solute** dissolved in the solvent

The molecules split up to form **hydrogen** ions.

A **strong** acid is **completely ionised** in aqueous solution. E.g. Hydrochloric, nitric and sulfuric acid.

A **weak** acid is only **partially ionised** in aqueous solution. E.g. Ethanoic, citric and carbonic.



A **weak acid** (aq) has a **lower pH** than a **strong acid** (aq) of the **same concentration**.

This is because a **weak acid** has a **lower concentration of hydrogen ions.**

As the pH decrease by one unit, the hydrogen ion concentration of the solution increase by a factor of 10.

Concentration of hydrogen ions in mol/dm^3	pH
0.10	1.0
0.010	2.0
0.0010	3.0
0.00010	4.0

Q1.

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

(a) Draw one line from each solution to the pH value of the solution.

Solution	**pH value of the solution**
	5
Acid	7
	9
Neutral	11
	13

(2)

(b) Which ion in aqueous solution causes acidity?

Tick **one** box.

H^+	☐
Na^+	☐
O^{2-}	☐
OH^-	☐

(1)

(c) When sulfuric acid is added to sodium hydroxide a reaction occurs to produce two products.

The equation is:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

How many elements are in the formula H_2SO_4?

Tick **one** box.

3 ☐

4 ☐

6 ☐

7 ☐

(1)

(d) What is this type of reaction?

Tick **one** box.

Decomposition ☐

Displacement ☐

Neutralisation ☐

Reduction ☐

(1)

(e) Name the salt produced.

____

(1)

(f) Describe how an indicator can be used to show when all the sodium hydroxide has reacted with sulfuric acid.

____

____

____

____

____

____

(3)

(Total 9 marks)

Q2.

A student plans a method to prepare pure crystals of copper sulfate.

The student's method is:

1. Add one spatula of calcium carbonate to dilute hydrochloric acid in a beaker.
2. When the fizzing stops, heat the solution with a Bunsen burner until all the liquid is gone.

The method contains several errors and does not produce copper sulfate crystals.

Explain the improvements the student should make to the method so that pure crystals of copper sulfate are produced.

(Total 6 marks)

Q3.

Tablets to cure indigestion contain a mixture that has been designed as a useful product.

(a) Complete the sentence.

Choose the answer from the box.

catalyst	formulation	hydrocarbon	solvent

Tablets to cure indigestion are an example of a ____________________ .

(1)

The table shows the substances in one tablet.

Substance	**Mass in mg**
Sodium hydrogencarbonate	64
Calcium carbonate	522
Magnesium carbonate	68

(b) The total mass of these substances in the tablet is 654 mg

What is the approximate fraction of magnesium carbonate in the total mass of these substances?

Tick **one** box.

$\frac{1}{10}$ ☐ $\frac{1}{100}$ ☐ $\frac{1}{1000}$ ☐ $\frac{1}{10\ 000}$ ☐

(1)

(c) The tablets also contain sugar.

Suggest why.

____

____

(1)

(d) Sodium hydrogencarbonate cures indigestion by reacting with acid in the stomach.

What type of reaction is this?

Tick **one** box.

Combustion ☐

Displacement ☐

Neutralisation ☐

(1)

A student adds an indigestion tablet to dilute hydrochloric acid.

(e) The gas produced is bubbled through limewater.

The gas turns the limewater milky.

Name the gas produced.

____

(1)

(f) Water is also produced.

Which **two** statements are reasons why water is a liquid at room temperature?

Tick **two** boxes.

Water has a boiling point of 100 °C	☐
Water has a giant covalent structure	☐
Water has a melting point lower than room temperature	☐
Water has delocalised electrons	☐
Water is made of ions	☐

(2)

(g) Calcium chloride is also produced.

- The formula for a calcium ion is Ca^{2+}
- The formula for a chloride ion is Cl^{-}

What is the formula of calcium chloride?

Tick **one** box.

$CaCl$ ☐ Ca_2Cl ☐ $CaCl_2$ ☐ Ca_2Cl_2 ☐

(1)

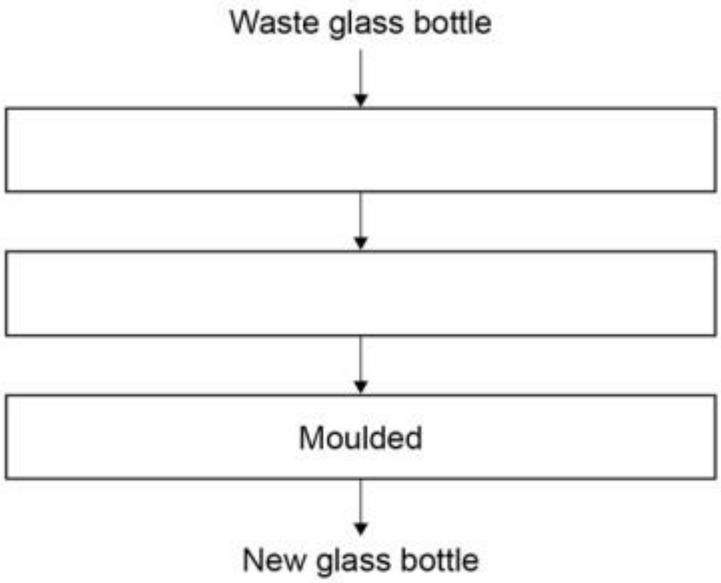
(h) The tablets are stored in glass bottles.

The flow chart shows part of a flowchart for recycling glass.

Complete the flow chart.

Choose the answers from the box.

crushed	**electrolysed**	**frozen**	**melted**	**oxidised**



(2)
(Total 10 marks)

Q4.

This question is about acids and bases.

(a) Which ion is found in all acids?

Tick **one** box.

Cl^- ☐ H^+ ☐ Na^+ ☐ OH^- ☐

(1)

(b) Zinc nitrate can be produced by reacting an acid and a metal oxide.

Name the acid and the metal oxide used to produce zinc nitrate.

Acid ____

Metal oxide ____

(2)

(c) In an equation, zinc nitrate is written as $Zn(NO_3)_2(aq)$.

What does (aq) mean?

Tick **one** box.

Dissolved in water ☐

Insoluble ☐

Not all reacted ☐

Reactant ☐

(1)

(d) The pH of a solution is 8

Some hydrochloric acid is added to the solution.

Suggest the pH of the solution after mixing.

pH = ___________________________

(1)

(e) **Table 1** shows the solubility of three solids in water at room temperature.

Table 1

Solid	**The mass of the solid that dissolves in 100 cm^3 of water**
Phosphorus oxide	50 g
Silicon dioxide	0 g
Sodium hydroxide	100 g

A teacher labelled these three solids **A**, **B** and **C**.

She gave a student the information shown in **Table 2**

Table 2

Solid	**Observation when added to water**	**pH of the solid in water**
A	colourless solution	14
B	colourless solution	2
C	solid does not dissolve	7

Describe a method that could be used to identify each of the three solids **A**, **B** and **C**.

You must use an indicator in the method.

Use information in **Table 1** and **Table 2**

____

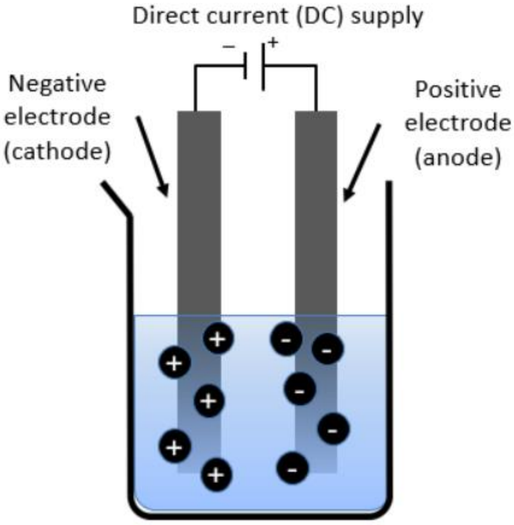
____

____

(4)
(Total 9 marks)

5.4.3 Electrolysis

When an **ionic compound** is **melted** or **dissolved in water**, the **ions** are **free** to **move** about the liquid or solution. These liquids and solutions are able to **conduct electricity** and are called **electrolytes**. Passing an **electric current** though electrolytes causes the ions to move to the electrodes.



Positive ions go to **negative** electrode (cathode) and are **reduced** (gain of electrons).

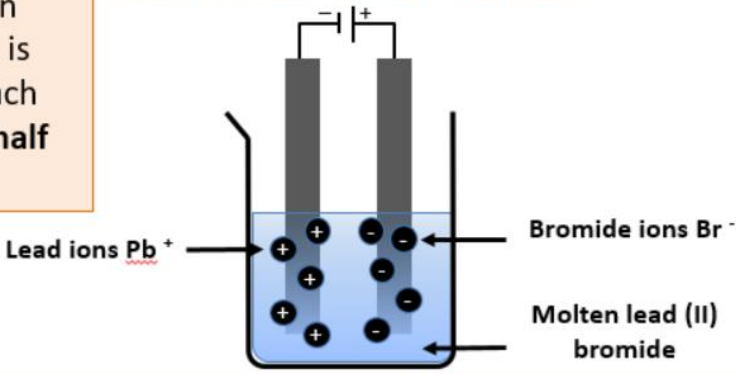
Negative ions go to the **positive** electrode (anode) and are **oxidised** (loss of electrons).

Ions are **discharged** at the electrodes producing **elements**. This is called **electrolysis**.

When an **ionic compound** is electrolysed in a **molten** state using inert electrodes, the **metal** is produced at the **cathode** and the **non-metal** is produced at the **anode**.

Higher: You can represent what is happening at each electrode using **half equations.**

lead bromide → lead + bromine



The positively charged lead ions Pb + (cations) are attracted to cathode and the negatively charged bromide ions Br $^-$ are attracted to the anode.

Higher:
At the cathode
$Pb^{2+} + 2e^- \rightarrow Pb$

Higher:
At the anode
$2Br^- \rightarrow Br_2 + 2e^-$ or $2Br^- - 2e^- \rightarrow Br_2$

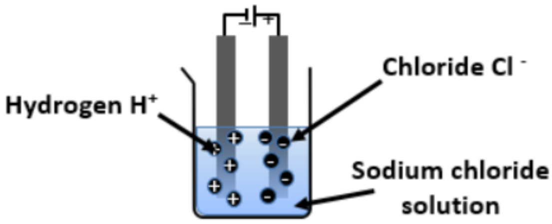
The **ions** discharged when an **aqueous solution** is electrolysed using inert electrodes depend on the relative **reactivity** of the elements involved.

At the **negative** electrode:
Metal will be produced on the electrode if it is **less** reactive than **hydrogen**. **Hydrogen** will be produced if the metal is **more** reactive than hydrogen.

At the **positive** electrode:
Oxygen is formed at positive electrode.

Higher: At the anode
$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
or $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$

If you have a **halide** ion (Cl^-, I^-, Br^-) then you will get **chlorine, bromine or iodine** formed at that electrode.
This happens because in the aqueous solution, **water molecules** break down producing **hydrogen** ions and **hydroxide** ions that are discharged.
$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$



sodium chloride → hydrogen + chlorine + sodium hydroxide

Uses of the products:
Chlorine: Bleach and PVC
Hydrogen: Margarine
Sodium hydroxide: Bleach and soap

Higher:
At the cathode
$2H^+ + 2e^- \rightarrow H_2$

Higher:
At the anode
$2Cl^- \rightarrow Cl_2 + 2e^-$ or $2Cl^- - 2e^- \rightarrow Cl_2$

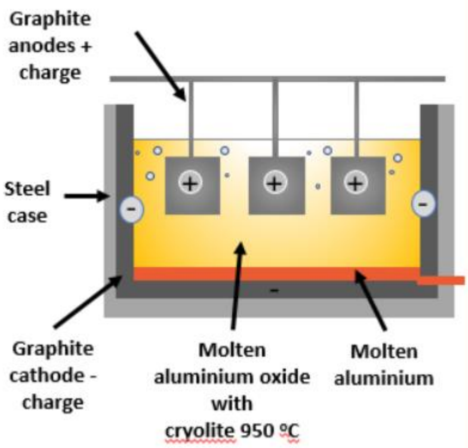
Metals can be extracted from **molten compounds** using electrolysis.

It is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.

Large amounts of **energy** are used in the extraction process to melt the compounds and to produce the electrical current.

Aluminum is manufactured by electrolysis of molten aluminum oxide.

Aluminium oxide → aluminium + oxygen



Aluminium oxide has a very **high melting point** so is mixed with molten **cryolite** to lower the temperature required to carry out the electrolysis.

Aluminium goes to the negative electrode and sinks to bottom.

Higher: $Al^{3+} + 3e^{-} \rightarrow Al$

Oxygen forms at positive electrodes.

Higher: $2O^{2-} \rightarrow O_2 + 4e^{-}$

The oxygen reacts with the carbon electrode making carbon dioxide causing damage. The electrode needs **replacing** due to this reaction.

$C + O_2 \rightarrow CO_2$

5.4.3 Electrolysis

EXAM QUESTIONS

Q1.

This question is about electrolysis.

(a) How many different elements are in the formula $AgNO_3$?

Tick **one** box.

2 ☐ 3 ☐ 5 ☐ 6 ☐

(1)

(b) How many atoms are in the formula $AgNO_3$?

Tick **one** box.

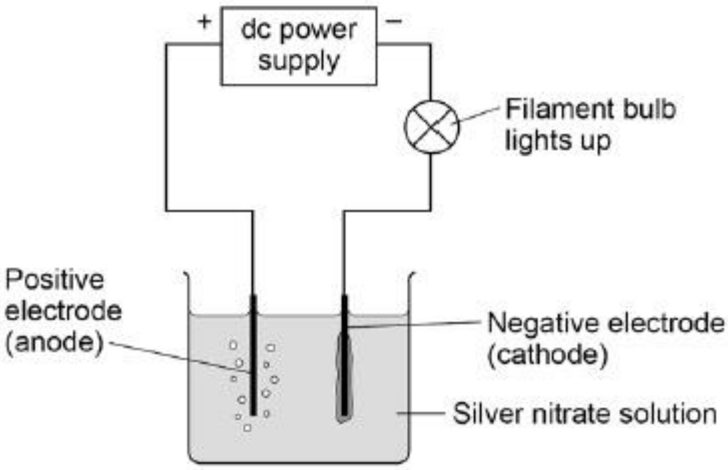
2 ☐ 3 ☐ 5 ☐ 6 ☐

(1)

An electric current is passed through silver nitrate solution.

Figure 1 shows the apparatus.

Figure 1



The solution contains four ions:

- Ag^+
- H^+
- NO_3^-
- OH^-

(c) Where do the H^+ and OH^- ions come from?

Tick **one** box.

Air	☐
Electrodes	☐
Silver nitrate	☐
Water	☐

(1)

(d) Ag^+ ions and H^+ ions are attracted to the negative electrode (cathode).

Give a reason why.

____

____

(1)

(e) Silver is produced at the negative electrode (cathode) and not hydrogen.

What does this tell you about the reactivity of silver?

Tick **one** box.

Silver is less reactive than hydrogen	☐
Silver is less reactive than oxygen	☐

Silver is more reactive than nitrate	☐
Silver is more reactive than water	☐

(1)

(f) The hydroxide ion (OH^-) is attracted to the positive electrode (anode).

The equation shows what happens at the positive electrode (anode).

$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$

Name the gas produced at the positive electrode (anode).

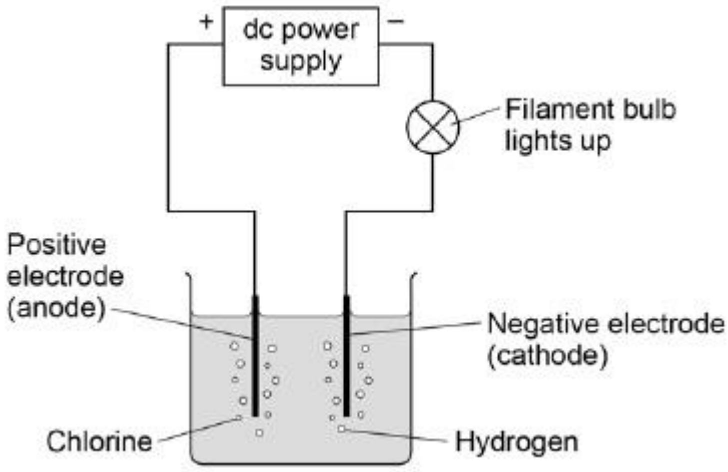
____

(1)

(g) An electric current is passed through sodium chloride solution.

Figure 2 shows the apparatus.

Figure 2



After passing an electric current through sodium chloride solution one product is sodium hydroxide (NaOH) solution.

The presence of sodium hydroxide can be shown by adding an indicator.

Name an indicator.

Give the colour of the indicator in sodium hydroxide solution.

Indicator ____

Colour ____

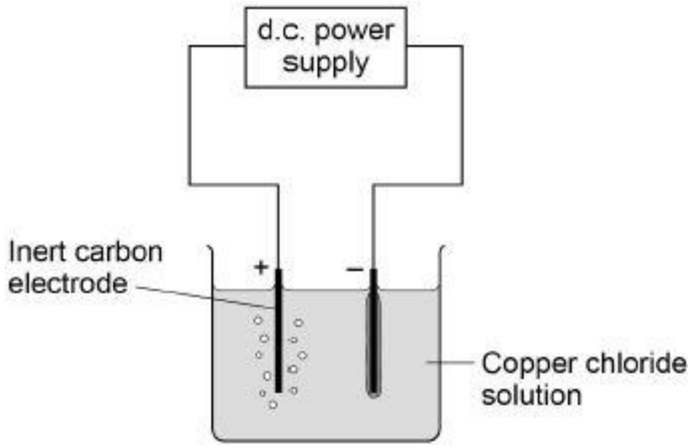
(2)
(Total 8 marks)

Q2.

This question is about electrolysis.

A student investigates the mass of copper produced during electrolysis of copper chloride solution.

The diagram below shows the apparatus.



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

Tick **one** box.

carbon dioxide ☐

chlorine ☐

hydrogen ☐

oxygen ☐

(1)

(b) Copper is produced at the negative electrode (cathode).

What does this tell you about the reactivity of copper?

Tick **one** box.

Copper is less reactive than hydrogen ☐

Copper is less reactive than oxygen ☐

Copper is more reactive than carbon ☐

Copper is more reactive than chlorine ☐

(1)

The table below shows the student's results.

	Total mass of copper produced in mg			
Time in mins	**Experiment 1**	**Experiment 2**	**Experiment 3**	**Mean**
1	0.60	0.58	0.62	0.60
2	1.17	1.22	1.21	1.20
4	2.40	2.41	2.39	2.40
5	3.02	**X**	3.01	3.06

(c) Determine the **mean** mass of copper produced after 3 minutes.

____

____

Mass = ______________________ mg

(1)

(d) Calculate the mass **X** of copper produced in **Experiment 2** after 5 minutes.

Use the table above.

____

____

____

____

Mass **X** = ______________________ mg

(2)

(e) The copper chloride solution used in the investigation contained 300 grams per dm^3 of solid $CuCl_2$ dissolved in 1 dm^3 of water.

The students used 50 cm^3 of copper chloride solution in each experiment.

Calculate the mass of solid copper chloride used in each experiment.

____

____

____

____

____

____

Mass = ______________________ g

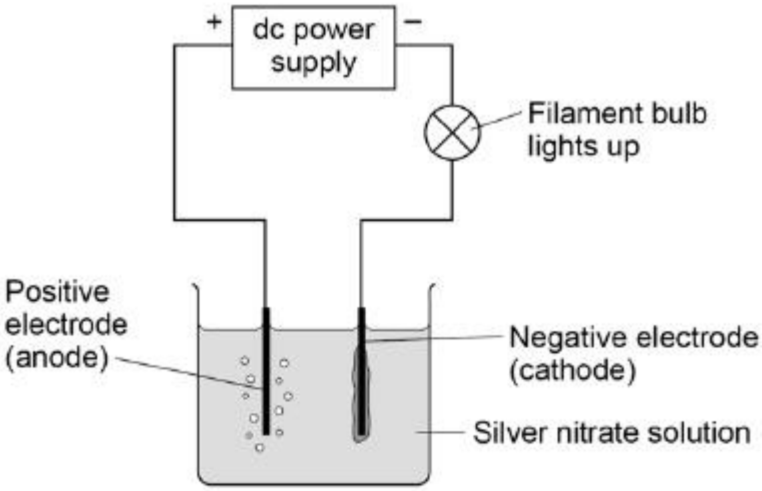
(3)(Total 8 marks)

Q3. (HT)

This question is about electrolysis.

(a) **Figure 1** shows the apparatus used to electrolyse silver nitrate ($AgNO_3$) solution.

Figure 1



Name the product discharged at each electrode.

Write a half equation for the reaction at each electrode.

Product at negative electrode (cathode) ____________________________________

Half equation for negative electrode

____

Product at positive electrode (anode) ______________________________________

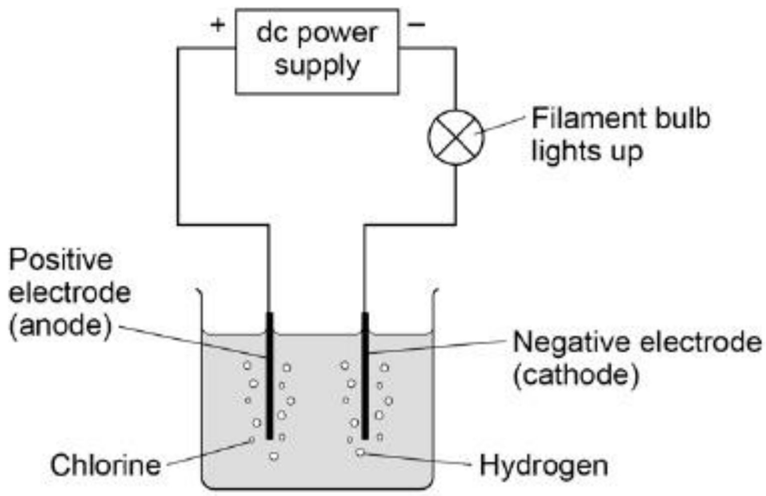
Half equation for positive electrode

____

(4)

(b) **Figure 2** shows the apparatus used to electrolyse sodium chloride (NaCl) solution.

Figure 2



Hydrogen and chlorine are produced.

Explain how another different product is formed in solution during this electrolysis.

____

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____

____

____

(4)
(Total 8 marks)

5.5.1 Endothermic and exothermic reactions

Energy is conserved in chemical reactions. The amount of energy in the Universe at the end of a chemical reaction is the same as before the reaction takes place.

$$H_2\,(g) + Cl_2\,(g) \rightarrow 2HCl\,(g)$$

In the above reaction energy is released, it gets hotter.

An exothermic reaction is one that transfers energy to the surroundings so the temperature of the surroundings increases – "it gets hotter".

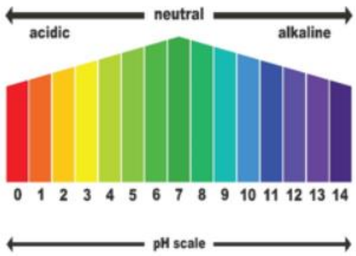
The two HCl molecules made will not hold as much energy as the H_2 and Cl_2 molecules at the start, so the spare energy is released as heat.

There are a number of common exothermic reactions, they include:

Combustion

Oxidation

Neutralisation



Know all three of these examples of exothermic reactions

Everyday uses of exothermic reactions include –
Self-heating cans
Hand warmers

Know both of these uses for exothermic reactions

We have already learnt that energy is conserved in chemical reactions.

$2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O\ (l)$

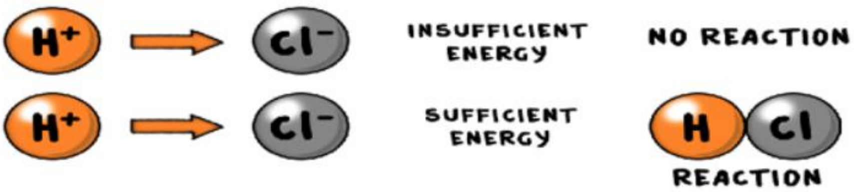
In the above reaction, energy is taken in- it gets colder.
An endothermic reaction is one that takes energy from the surroundings so the temperature of the surroundings decreases – "it gets colder".
The sodium ethanoate, carbon dioxide and water molecules made will hold more energy than the ethanoic acid and sodium carbonate molecules at the start, so the energy needed is taken in as heat.

Know all three of these examples of endothermic reactions

Other examples of endothermic reactions are

- Thermal decomposition
- Sports injury packs

Chemical reactions can only occur when reacting particles collide with each other with sufficient energy.



The minimum amount of energy that particles must have to react is called the activation energy

You have given a reaction its activation energy when you have used a lit spill to light a Bunsen burner. Without the activation energy from the lit spill the methane gas and oxygen in the air will not combust and release the heat energy.

You will be expected to balance this equation.

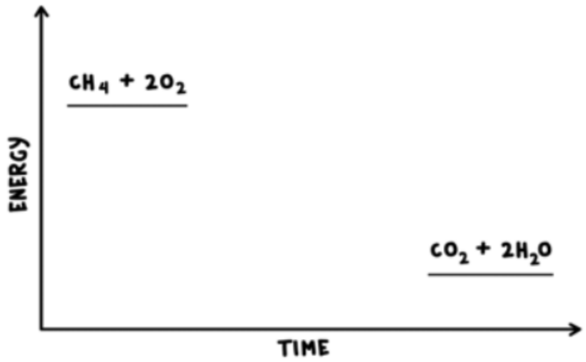
When we look at this reaction we see the following.

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

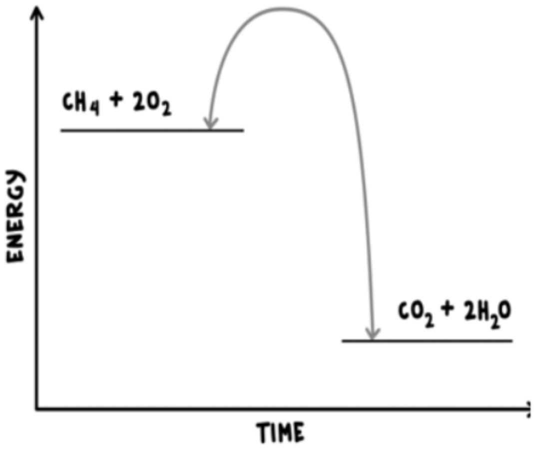
$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

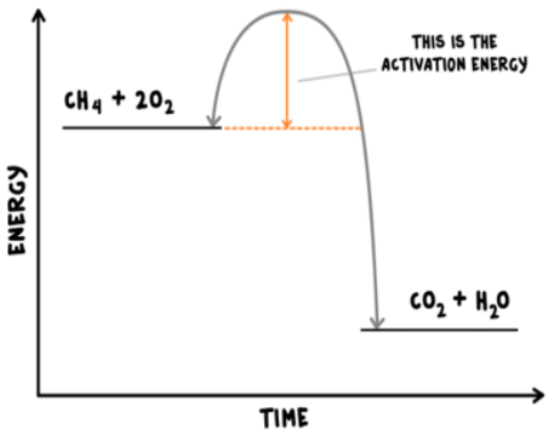
We know this reaction is exothermic, this means energy is released. So the CH_4 and $2O_2$, the reactants, must have more energy than the products, CO_2 and $2H_2O$

We can show this as a reaction profile. On it we need to include the formulae or names of the products and reactants. We also need to show the relative energies of the reactants and products

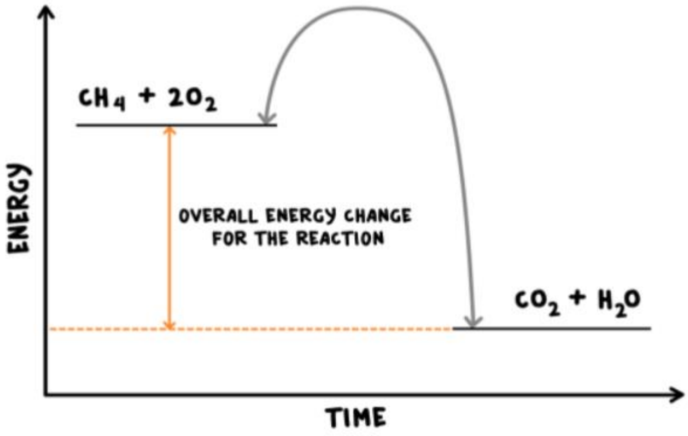


More information needs to be included in the reaction profile. This will show the activation energy of the reaction. It is shown by a curved line rising above the reactants energy.





We can now see the overall change in energy within the reaction.

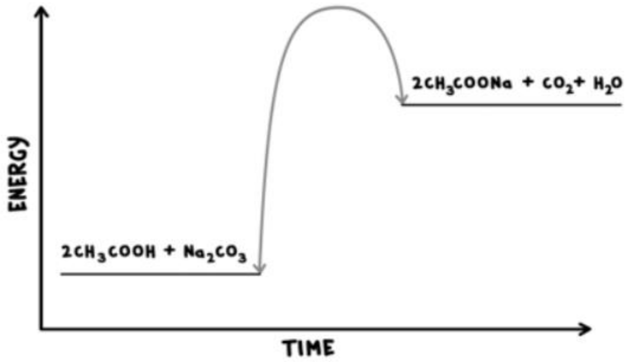


The products have less energy than the reactants. This will have been lost as heat as the reaction is exothermic.

We saw earlier that the following reaction was endothermic:

$$2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O\ (l)$$

What would the reaction profile look like for this reaction?



We can see that the products have more energy than the reactants. This will have been taken in as heat energy it feels colder.

EXAM QUESTIONS

Q1.

This question is about compounds of oxygen.

The reaction between carbon and oxygen is exothermic.

(a) What does exothermic reaction mean?

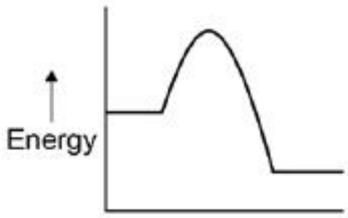
____

____

(1)

(b) Which is the correct reaction profile (energy level diagram) for an exothermic reaction?

Tick **one** box.



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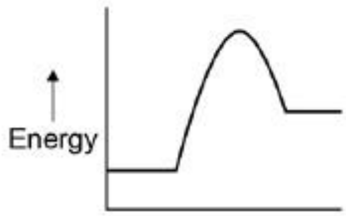

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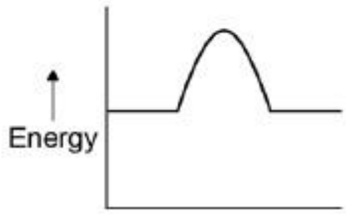


☐

(1)

(c) The percentage by mass of oxygen in carbon dioxide (CO_2) is calculated by the equation:

$$\text{percentage by mass} = \frac{\text{number of atoms of O} \times \text{Relative atomic mass of oxygen (O)}}{\text{relative molecular mass of carbon dioxide } (CO_2)} \times 100$$

Relative atomic masses (A_r): C = 12 O = 16

Calculate the percentage by mass of oxygen in carbon dioxide (CO_2).

____

____

____

____

____

____

____

Percentage by mass of oxygen = ____________________ %

(3)

Hydrogen peroxide decomposes to produce water and oxygen.

(d) Balance the chemical equation.

$$\_\_\_\_ \; H_2O_2 \rightarrow \_\_\_\_ \; H_2O + O_2$$

(1)

(e) 6.8 g of hydrogen peroxide decomposes to produce 3.6 g of water.

Calculate the mass of oxygen produced when 68 g of hydrogen peroxide decomposes.

____

____

____

____

Mass of oxygen = ____________________ g

(2)

(Total 8 marks)

Q2.

Some students investigated the change in temperature as sodium hydroxide solution is added to dilute sulfuric acid.

This is the method used.

1. Put 25 cm^3 of dilute sulfuric acid into a polystyrene cup.

2. Measure the initial temperature of the dilute sulfuric acid.

3. Add 4 cm^3 of sodium hydroxide solution to the dilute sulfuric acid.

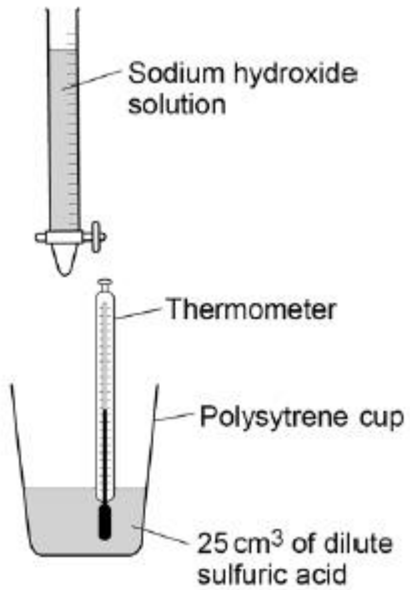
4. Stir the mixture.

5. Measure the highest temperature of the mixture.

6. Repeat steps 3–5 until 40 cm^3 of sodium hydroxide solution have been added.

Figure 1 shows the apparatus the student used.

Figure 1



(a) The volume of sodium hydroxide solution is a variable.

Which **two** words can be used to describe this type of variable?

Tick **two** boxes.

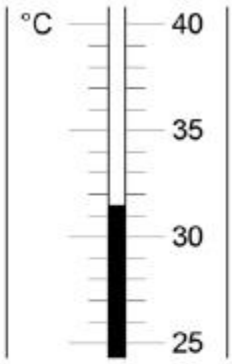
Categoric	☐
Continuous	☐
Control	☐
Dependent	☐
Independent	☐

(2)

(b) The dilute sulfuric acid has an initial temperature of 24.0 °C

Figure 2 shows the highest temperature.

Figure 2



Calculate the change in temperature.

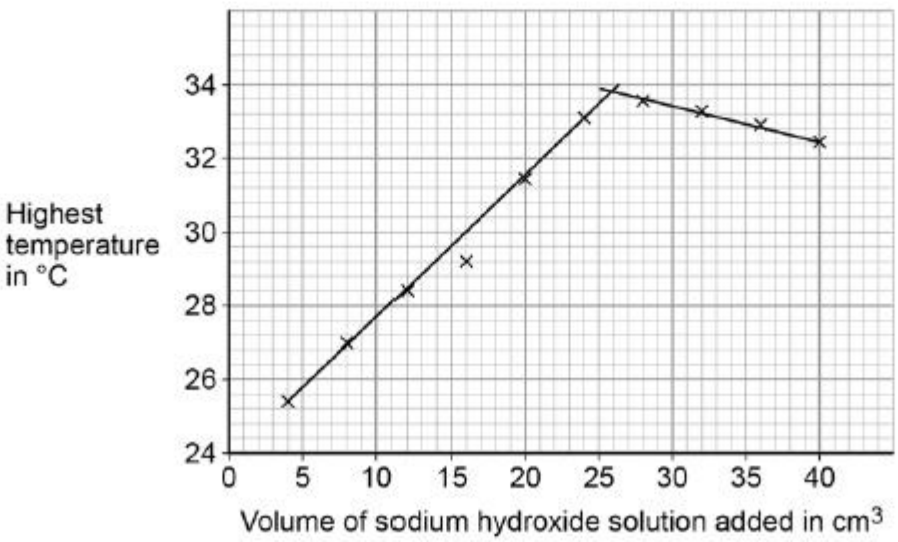
____

Temperature = ____________________ °C

(2)

Figure 3 shows the students' results.

Figure 3



(c) Determine the volume of sodium hydroxide solution that gives the highest temperature change.

Use **Figure 3** to help you answer this question.

Volume = ____________________ cm^3

(1)

(d) In **Figure 3** the temperature when 16 cm^3 of sodium hydroxide solution is added is anomalous.

Suggest **one** error that could have been made in the method which would cause this anomalous result.

____

____

(1)

(e) The sodium hydroxide solution in this investigation contains 80 grams per dm^3

The students use 40 cm^3 of sodium hydroxide solution.

Calculate the mass of sodium hydroxide in 40 cm^3

____

____

____

____

____

Mass = ____________________ g

(3)

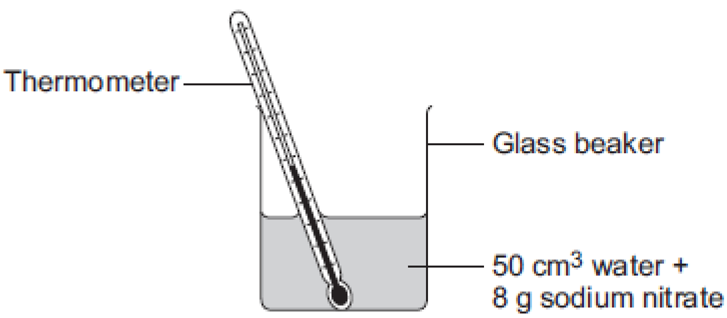
(Total 9 marks)

Q3.

This question is about temperature changes.

(a) A student investigated the temperature change when 8 g of sodium nitrate dissolves in 50 cm^3 of water.

The diagram below shows the apparatus the student used.



The student did the experiment five times.
Table 1 shows the results.

Table 1

Experiment	Decrease in temperature of water in °C
1	5.9
2	5.7
3	7.2
4	5.6
5	5.8

(i) Calculate the mean decrease in temperature.
Do not use the anomalous result in your calculation.

___

___

Mean decrease in temperature = ____________ °C

(2)

(ii) Suggest **one** change in the apparatus in the diagram above which would improve the accuracy of the results.
Give a reason for your answer.

___

___

___

___

(2)

(b) The student investigated the temperature change when different masses of sodium carbonate were added to 50 cm^3 of water at 20 °C.

Table 2 below shows the results.

Table 2

Mass of sodium carbonate in g	Final temperature of solution in °C
2.0	21.5
4.0	23.0
6.0	24.5
8.0	26.0

10.0	26.6
12.0	26.6
14.0	26.6

Describe the relationship between the mass of sodium carbonate added and the final temperature of the solution.

Use values from **Table 2** in your answer.

___

___

___

___

___

___

(3)
(Total 7 marks)

HT ONLY

During a chemical reaction:

- **Energy must be supplied to break bonds in the reactants**
- **Energy is released to form bonds in the products.**

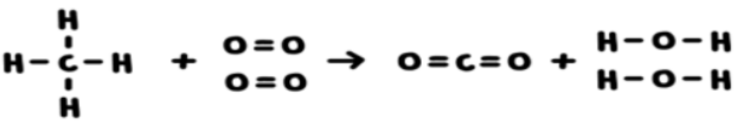
The energy needed to break bonds and the energy released when bonds are formed can be calculated from bond energies.

Bond	Bond Energy kJ/mol
C-H	411
O=O	494
C=O	799
O-H	459

This means that 411kJ/mol of energy needs to be put in to break the carbon-hydrogen bond.
It also means that 459 kJ/mol is given out when the oxygen hydrogen bond is made in water.

The difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction.

Worked example



For the reactants	For the products
There are four C-H bonds so 4 x 411kJ/mol = 1,644kJ/mol	There are two C=O bonds so 2 x 799kJ/mol = 1,598kJ/mol
There are two O=O bonds so 2 x 494kJ/mol = 988kJ/mol	There are four O-H bonds so 4 x 459kJ/mol = 1,836kJ/mol
The sum of these is the energy supplied to break the bonds in the reactants it is 1,644 + 988 = 2,632kJ/mol	The sum of these is the energy released when bonds in the products are formed it is 1,598 + 1,836 = 3,434kJ/mol

We already know that the difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction.

This means that

Overall energy change = energy needed to break the bonds – energy released as bonds are made

= 2,632kJ/mol - 3,434kJ/mol

Overall energy change = - 802kJ/mol

This is an exothermic reaction, so the sum of the difference between the calculations is negative. For an endothermic reaction it would be positive.

Students should be able to calculate the energy transferred in chemical reactions using bond energies supplied.

Know these two definitions- they are often asked for in the exam.

In an **exothermic** reaction, the energy released from forming **new bonds is greater** than the energy needed to break existing bonds

In an **endothermic** reaction, the energy needed to **break existing bonds is greater** then the energy released from forming new bonds

EXAM QUESTION (HT)

Q1.

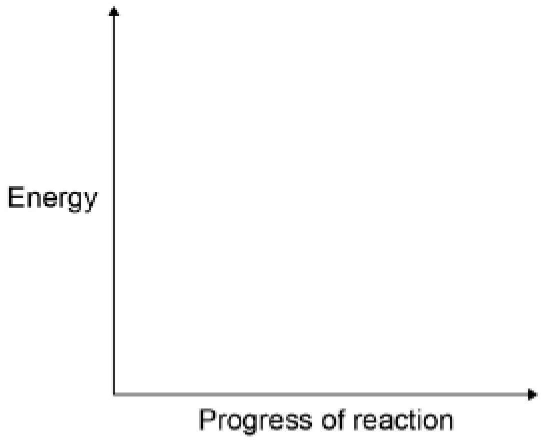
Exothermic reactions transfer energy to the surroundings.

(a) Draw a reaction profile for an exothermic reaction using the axes in **Figure 1**.

Show the:

- relative energies of the reactants and products
- activation energy and overall energy change.

Figure 1

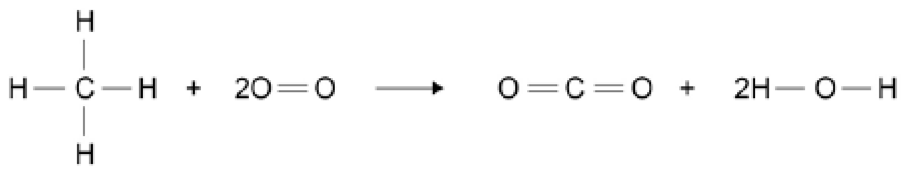


(2)

(b) Combustion is an exothermic reaction.

Calculate the overall energy change for the complete combustion of one mole of methane in oxygen.

$$CH_4 \quad + \quad 2O_2 \quad \longrightarrow \quad CO_2 \quad + \quad 2H_2O$$



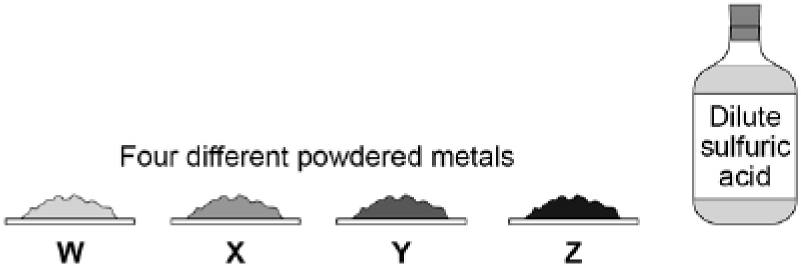
Bond	Bond energy in kJ / mol
C — H	413
O ═ O	498
C ═ O	805
O — H	464

Overall energy change = ________________ kJ / mol

(3)

(c) **Figure 2** shows the chemicals given to a student.

Figure 2



The student wants to investigate the reactivity of the four metals.

Outline a plan the student could use to investigate the relative reactivity of the four metals, **W**, **X**, **Y** and **Z**.

The plan should use the fact that all four metals react exothermically with dilute sulfuric acid.

You should name the apparatus used and comment on the safe use of the chemicals.

____

____

____

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____

(6)

(d) Another student used displacement reactions to investigate the relative reactivity of the four metals, **W**, **X**, **Y** and **Z**.

The table below shows the student's results.

	Observations			
Solution	**Metal W**	**Metal X**	**Metal Y**	**Metal Z**
Copper nitrate	Brown layer formed on metal	Brown layer formed on metal	Brown layer formed on metal	No change
Magnesium sulfate	No change	No change	No change	No change
Sulfuric acid	Gas bubbles produced	Few gas bubbles produced	Gas bubbles produced	No change
Zinc chloride	Grey layer formed on metal	No change	No change	No change

Give the order of reactivity of metals, **W**, **X**, **Y** and **Z**.

Use the results in the table above to justify your answer.

> > >

(3)

(e) The student concluded that these results could also be used to justify the order of reactivity of copper, magnesium, hydrogen and zinc.

The student is **not completely** correct. Use the results in the table above to explain why.

Suggest one further experiment that would provide evidence for the student's conclusion.

(4)
(Total 18 marks)

Q2.

This question is about energy changes in chemical reactions.

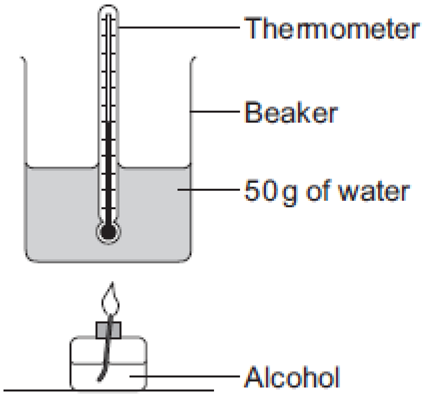
(a) Balance the chemical equation for the combustion of methane.

$$CH_4 \quad + \quad O_2 \quad \longrightarrow \quad CO_2 \quad + \quad H_2O$$

(1)

(b) Alcohols are used as fuels.

A group of students investigated the amount of energy released when an alcohol was burned. The students used the apparatus shown in the diagram below.



In one experiment the temperature of 50 g of water increased from 22.0 °C to 38.4 °C.
The mass of alcohol burned was 0.8 g.

Calculate the heat energy (Q) in joules, released by burning 0.8 g of the alcohol.
Use the equation:

$$Q = m \times c \times \Delta T$$

Specific heat capacity (c) = 4.2 J / g / °C

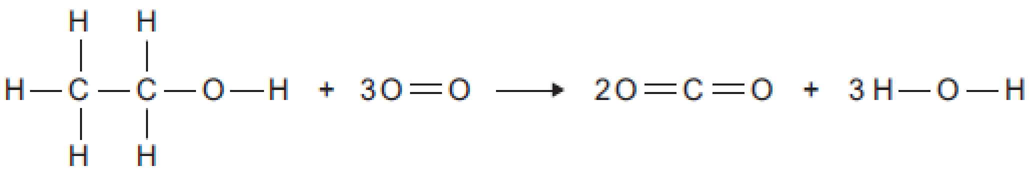
Heat energy (Q) = ____________________ J

(3)

(c) The chemical equation for the combustion of ethanol is:

$$C_2H_5OH \quad + \quad 3O_2 \quad \longrightarrow \quad 2CO_2 \quad + \quad 3H_2O$$

(i) The equation for the reaction can be shown as:



Bond	Bond energy in kJ per mole
C — H	413
C — C	347
C — O	358
C ═ O	799
O — H	467
O ═ O	495

Use the bond energies to calculate the overall energy change for this reaction.

____

____

____

____

____

____

Overall energy change = ____________ kJ per mole

(3)

(ii) The reaction is exothermic.
Explain why, in terms of bonds broken and bonds formed.

____

____

____

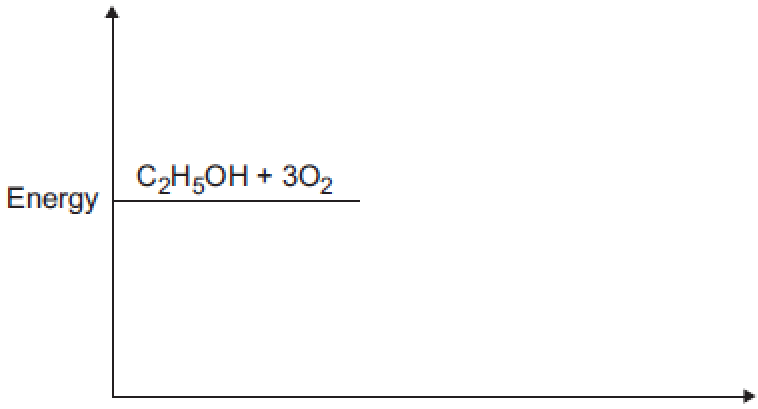
____

(2)

(iii) Complete the energy level diagram for the combustion of ethanol.

On the completed diagram, label:

- activation energy
- overall energy change.



(3)
(Total 12 marks)