AQA GCSE CHEMISTRY (HIGHER) · 8462

CHEMISTRY



EXAM QUESTION PACK

4.3.2 USE OF AMOUNT OF SUBSTANCE IN RELATION TO MASSES OF PURE SUBSTANCES

SCORE	PERCENTAGE
/208	%

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Q1.

This question is about elements in Group 1.

A teacher burns sodium in oxygen.

(a) Complete the word equation for the reaction.

sodium + oxygen ⟶ _____

(1)

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

Tick **one** box.

Decomposition ☐

Electrolysis ☐

Oxidation ☐

Precipitation ☐

(1)

(c) The teacher dissolves the product of the reaction in water and adds universal indicator.

The universal indicator turns purple.

What is the pH value of the solution?

Tick **one** box.

1	☐	4	☐	7	☐	13	☐

(1)

(d) The solution contains a substance with the formula NaOH

Give the name of the substance.

_____

(1)

(e) All alkalis contain the same ion.

What is the formula of this ion?

Tick **one** box.

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H^+ ☐

Na^+ ☐

OH^- ☐

O^{2-} ☐

(1)

(f) A solution of NaOH had a concentration of 40 g/dm^3

What mass of NaOH would there be in 250 cm^3 of the solution?

_____

_____

_____

_____

Mass = _____ g

(2)

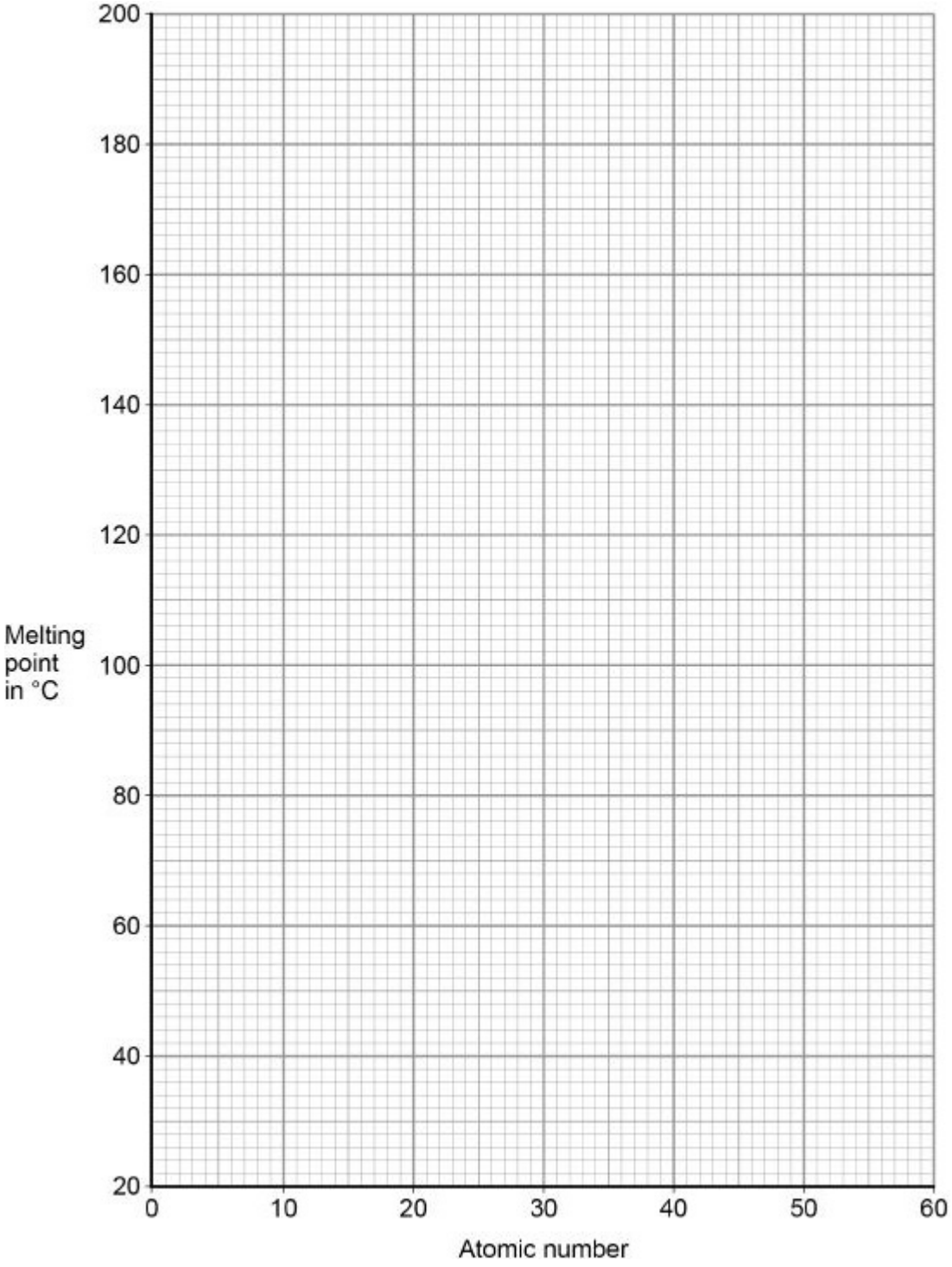
(g) The melting points of the elements in Group 1 show a trend.

The table below shows the atomic numbers and melting points of the Group 1 elements.

Element	**Atomic number**	**Melting point in °C**
Lithium	3	181
Sodium	11	98
Potassium	19	63
Rubidium	37	**X**
Caesium	55	29

Plot the data from the table on the graph below.

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(2)

(h) Predict the melting point, **X**, of rubidium, atomic number 37

Use the graph above.

Melting point = _____ °C

(1)

(Total 10 marks)

Q2.

Titanium is a transition metal.

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Titanium is extracted from titanium dioxide in a two-stage industrial process.

Stage 1 $TiO_2 + 2\ C + 2\ Cl_2 \longrightarrow TiCl_4 + 2\ CO$

Stage 2 $TiCl_4 + 4\ Na \longrightarrow Ti + 4\ NaCl$

(a) Suggest **one** hazard associated with **Stage 1**.

_____

_____

(1)

(b) Water must be kept away from the reaction in **Stage 2**.

Give **one** reason why it would be hazardous if water came into contact with sodium.

_____

_____

(1)

(c) Suggest why the reaction in **Stage 2** is carried out in an atmosphere of argon and **not** in air.

_____

_____

_____

_____

(2)

(d) Titanium chloride is a liquid at room temperature.

Explain why you would **not** expect titanium chloride to be a liquid at room temperature.

_____

_____

_____

_____

_____

_____

_____

(3)

In **Stage 2**, sodium displaces titanium from titanium chloride.

(e) Sodium atoms are oxidised to sodium ions in this reaction.

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Why is this an oxidation reaction?

_____

_____

(1)

(f) Complete the half equation for the oxidation reaction.

$$Na \longrightarrow _____ + _____$$

(1)

(g) In Stage 2, 40 kg of titanium chloride was added to 20 kg of sodium.

The equation for the reaction is:

$$TiCl_4 + 4\ Na \longrightarrow Ti + 4\ NaCl$$

Relative atomic masses (A_r): Na = 23 Cl = 35.5 Ti = 48

Explain why titanium chloride is the limiting reactant.

You **must** show your working.

_____

_____

_____

_____

_____

_____

_____

_____

(4)

(h) For a **Stage 2** reaction the percentage yield was 92.3%

The theoretical maximum mass of titanium produced in this batch was 13.5 kg.

Calculate the actual mass of titanium produced.

_____

_____

_____

Mass of titanium = _____ kg

(2)

(Total 15 marks)

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Q3.

This question is about acids and alkalis.

(a) Dilute hydrochloric acid is a strong acid.

Explain why an acid can be described as both strong and dilute.

_____

_____

_____

_____

(2)

(b) A 1.0×10^{-3} mol/dm^3 solution of hydrochloric acid has a pH of 3.0

What is the pH of a 1.0×10^{-5} mol/dm^3 solution of hydrochloric acid?

pH = _____

(1)

A student titrated 25.0 cm^3 portions of dilute sulfuric acid with a 0.105 mol/dm^3 sodium hydroxide solution.

(c) The table below shows the student's results.

	Titration 1	**Titration 2**	**Titration 3**	**Titration 4**	**Titration 5**
Volume of sodium hydroxide solution in cm^3	23.50	21.10	22.10	22.15	22.15

The equation for the reaction is:

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

Calculate the concentration of the sulfuric acid in mol/dm^3

Use only the student's concordant results.

Concordant results are those within 0.10 cm^3 of each other.

_____

_____

_____

_____

_____

_____

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_____

_____

_____

_____

Concentration of sulfuric acid = _____ mol/dm^3

(5)

(d) Explain why the student should use a pipette to measure the dilute sulfuric acid and a burette to measure the sodium hydroxide solution.

_____

_____

_____

_____

(2)

(e) Calculate the mass of sodium hydroxide in 30.0 cm^3 of a 0.105 mol/dm^3 solution.

Relative formula mass (M_r): NaOH = 40

_____

_____

_____

_____

Mass of sodium hydroxide = _____ g

(2)

(Total 12 marks)

Q4.

This question is about methanol.

(a) Methanol is broken down in the body during digestion.

What type of substance acts as a catalyst in this process?

Tick **one** box.

Amino acid ☐

Enzyme ☐

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Ester ☐

Nucleotide ☐

(1)

In industry, methanol is produced by reacting carbon monoxide with hydrogen.

The equation for the reaction is:

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

(b) How many moles of carbon monoxide react completely with 4.0×10^3 moles of hydrogen?

Tick **one** box.

1.0×10^3 moles ☐

2.0×10^3 moles ☐

4.0×10^3 moles ☐

8.0×10^3 moles ☐

(1)

(c) The reaction is carried out at a temperature of 250 °C and a pressure of 100 atmospheres.

The forward reaction is exothermic.

Explain what happens to the yield of methanol if a temperature higher than 250 °C is used.

_____

_____

_____

_____

(2)

(d) A pressure of 100 atmospheres is used instead of atmospheric pressure.

The higher pressure gives a greater yield of methanol and an increased rate of reaction.

Explain why.

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_____

_____

_____

_____

_____

_____

_____

_____

_____

_____

_____

(4)

A catalyst is used in the reaction to produce methanol from carbon monoxide and hydrogen.

(e) Explain how a catalyst increases the rate of a reaction.

_____

_____

_____

_____

(2)

(f) Suggest why a catalyst is used in this industrial process.

Do **not** give answers in terms of increasing the rate of reaction.

_____

_____

(1)

(g) Suggest the effect of using the catalyst on the equilibrium yield of methanol.

_____

_____

(1)

(Total 12 marks)

Q5.

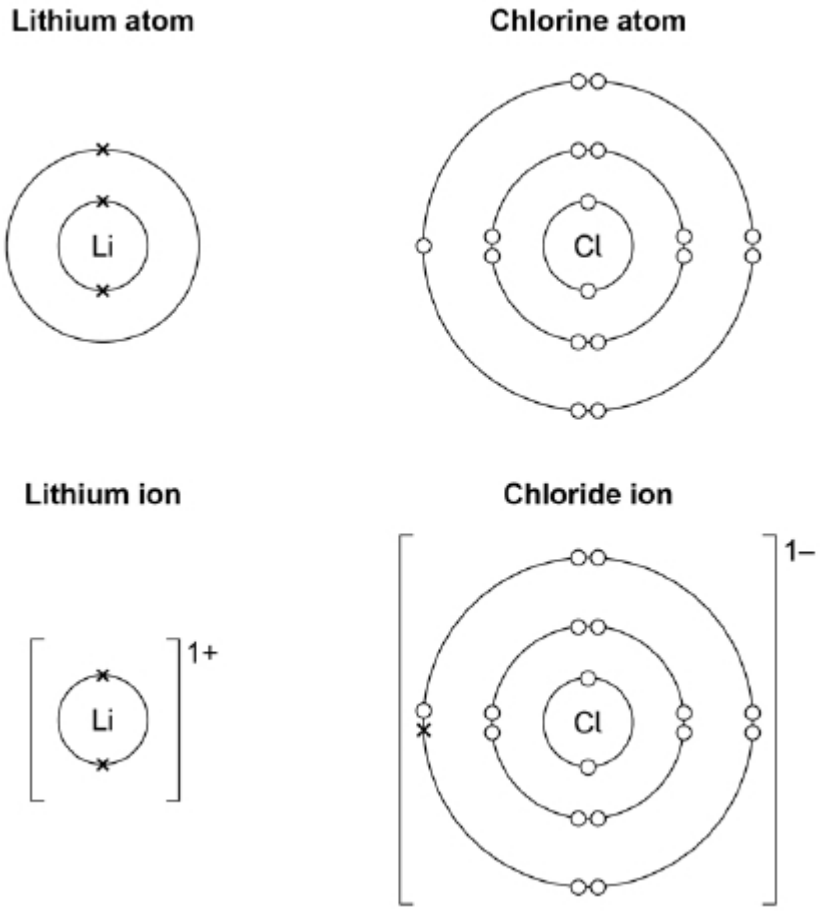
This question is about metal compounds.

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(a) Lithium reacts with chlorine to produce lithium chloride.

When lithium atoms and chlorine atoms react to produce lithium chloride, lithium ions and chloride ions are formed.

The diagram shows the electronic structures of the atoms and ions.
The symbols **o** and **x** are used to represent electrons.



Describe what happens when a lithium atom reacts with a chlorine atom.

Answer in terms of electrons.

_____

_____

_____

_____

_____

_____

_____

_____

(4)

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Zinc sulfate can be made by two methods.

The equations for the two methods are:

Method 1: $ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$

Method 2: $ZnCO_3 + H_2SO_4 \longrightarrow ZnSO_4 + H_2O + CO_2$

(b) Calculate the percentage atom economy for making zinc sulfate in **Method 1**.

Use the equation:

percentage atom economy =

$$\frac{\text{relative formula mass of } ZnSO_4}{\text{relative formula mass of } ZnO + \text{relative formula mass of } H_2SO_4} \times 100$$

Give your answer to 3 significant figures.

Relative formula masses (M_r): ZnO = 81 H_2SO_4 = 98 $ZnSO_4$ = 161

_____

_____

_____

_____

_____

Percentage atom economy = _____ %

(3)

(c) **Method 1** gives a higher percentage atom economy for making zinc sulfate than **Method 2**.

Give a reason why it is important to use a reaction with a high atom economy.

_____

_____

(1)

(d) A student uses 50 cm^3 of a zinc sulfate solution of 80 g/dm^3

What mass of zinc sulfate is dissolved in 50 cm^3 of this zinc sulfate solution?

_____

_____

_____

_____

Mass = _____ g

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(2)
(Total 10 marks)

Q6.

This question is about atoms.

(a) What does the number 19 represent in $^{19}_{9}F$?

_____

(1)

(b) How many atoms are present in one mole of fluorine atoms?

Tick (✔) **one** box.

2.03×10^{26} ☐

2.06×10^{23} ☐

6.02×10^{23} ☐

6.02×10^{26} ☐

(1)

(c) The plum pudding model of the atom was replaced by the nuclear model.

The nuclear model was developed after the alpha particle scattering experiment.

Compare the plum pudding model with the nuclear model of the atom.

_____

_____

_____

_____

_____

_____

_____

_____

(4)

(d) An element has three isotopes.

The table shows the mass numbers and percentage of each isotope.

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	Isotope 1	Isotope 2	Isotope 3
Mass number	24	25	26
Percentage (%)	78.6	10.1	11.3

Calculate the relative atomic mass (A_r) of the element.

Give your answer to 3 significant figures.

Relative atomic mass = _____

(2)

(Total 8 marks)

Q7.

A scientist produces zinc iodide (ZnI_2).

This is the method used.

1. Weigh 0.500 g of iodine.
2. Dissolve the iodine in ethanol.
3. Add an excess of zinc.
4. Stir the mixture until there is no further change.
5. Filter off the excess zinc.
6. Evaporate off the ethanol.

(a) Ethanol is flammable.

Suggest how the scientist could carry out **Step 6** safely.

(1)

(b) Explain why the scientist adds excess zinc rather than excess iodine.

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_____

(3)

(c) Calculate the minimum mass of zinc that needs to be added to 0.500 g of iodine so that the iodine fully reacts.

The equation for the reaction is:

$$Zn + I_2 \longrightarrow ZnI_2$$

Relative atomic masses (M_r): Zn = 65 I = 127

_____

_____

_____

_____

_____

_____

Minimum mass of zinc = _____ g

(3)

A different scientist makes zinc iodide by the same method.

The scientist obtains 12.5 g of zinc iodide.

The percentage yield in this reaction is 92.0%.

(d) What is the maximum theoretical mass of zinc iodide produced in this reaction?

_____

_____

_____

_____

_____

_____

Maximum theoretical mass = _____ g

(3)

(e) Suggest **one** reason why the percentage yield in this reaction is **not** 100%.

_____

_____

(1)

(f) The scientist makes a solution of zinc iodide with a concentration of 0.100 mol / dm^3

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Calculate the mass of zinc iodide (ZnI_2) required to make 250 cm^3 of this solution.

Relative atomic masses (A_r): Zn = 65 I = 127

_____

_____

_____

_____

_____

_____

Mass = _____ g

(3)

(Total 14 marks)

Q8.

Potable water is water that is safe to drink.

Seawater can be changed into potable water by desalination.

(a) Name the substance removed from seawater by desalination.

_____

(1)

(b) Desalination requires large amounts of energy.

Desalination is only used when there is no other source of potable water.

Give **one** reason why.

_____

_____

(1)

Water from lakes and rivers can be treated to make it potable.

(c) The first stage is to filter the water from lakes and rivers.

Why is the water filtered?

_____

_____

(1)

(d) Chlorine gas is then added to the filtered water.

Why is chlorine gas used to treat water?

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_____

_____

(1)

(e) Describe a test for chlorine gas.

Give the result of the test if chlorine is present.

Test _____

Result _____

(2)

Some students investigated different water samples.

The table shows some of their results.

Water	**pH**	**Mass of dissolved solid in g / dm^3**
Tap water	6.5	0.5
Seawater	8.1	35.0
Pure water		

(f) Complete the table above to show the expected results for pure water.

(2)

(g) What mass of dissolved solid is present in 100 cm^3 of the sample of tap water?

Tick (✔) **one** box.

0.05 g ☐

0.5 g ☐

5 g ☐

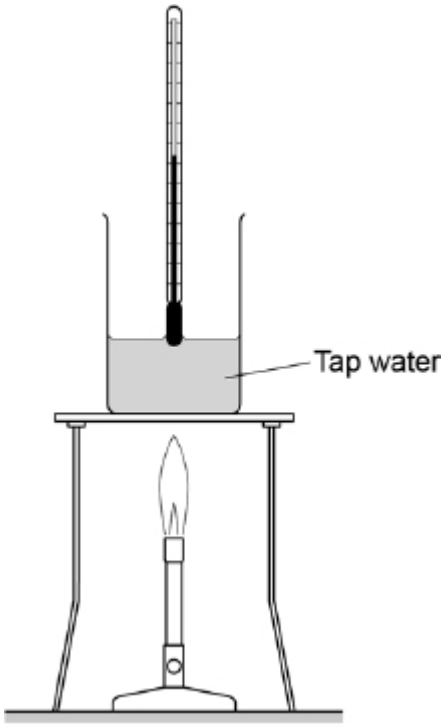
50 g ☐

(1)

(h) Boiling points can be used to show whether substances are pure.

The diagram shows the apparatus the students used to find the boiling point of tap water.

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The students made a mistake setting up the apparatus.

What mistake did the students make?

_____

_____

(1)

(Total 10 marks)

Q9.

A student investigated the reactions of copper carbonate and copper oxide with dilute hydrochloric acid.

In both reactions one of the products is copper chloride.

(a) Describe how a sample of copper chloride crystals could be made from copper carbonate and dilute hydrochloric acid.

_____

_____

_____

_____

_____

_____

_____

_____

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(4)

(b) A student wanted to make 11.0 g of copper chloride.

The equation for the reaction is:

$$CuCO_3 + 2HCl \rightarrow CuCl_2 + H_2O + CO_2$$

Relative atomic masses, A_r: H = 1; C = 12; O = 16; Cl = 35.5; Cu = 63.5

Calculate the mass of copper carbonate the student should react with dilute hydrochloric acid to make 11.0 g of copper chloride.

_____

_____

_____

_____

Mass of copper carbonate = _____ g

(4)

(c) The percentage yield of copper chloride was 79.1 %.

Calculate the mass of copper chloride the student actually produced.

_____

_____

Actual mass of copper chloride produced = _____ g

(2)

(d) Look at the equations for the two reactions:

Reaction 1 $CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$

Reaction 2 $CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$

Reactive formula masses: CuO = 79.5; HCl = 36.5; $CuCl_2$ = 134.5; H_2O = 18

The percentage atom economy for a reaction is calculated using:

$$\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}} \times 100$$

Calculate the percentage atom economy for Reaction 2.

_____

_____

_____

_____

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_____

_____

Percentage atom economy = _____ %

(3)

(e) The atom economy for Reaction 1 is 68.45 %.
Compare the atom economies of the two reactions for making copper chloride.

Give a reason for the difference.

_____

_____

(1)

(Total 14 marks)

Q10.

Sodium carbonate reacts with dilute hydrochloric acid:

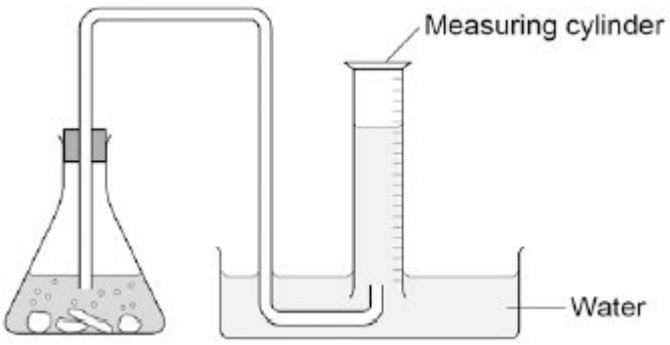
$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

A student investigated the volume of carbon dioxide produced when different masses of sodium carbonate were reacted with dilute hydrochloric acid.

This is the method used.

1. Place a known mass of sodium carbonate in a conical flask.
2. Measure 10 cm^3 of dilute hydrochloric acid using a measuring cylinder.
3. Pour the acid into the conical flask.
4. Place a bung in the flask and collect the gas until the reaction is complete.

(a) The student set up the apparatus as shown in the figure below.



Identify the error in the way the student set up the apparatus.

Describe what would happen if the student used the apparatus shown.

_____

_____

_____

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_____

(2)

(b) The student corrected the error.

The student's results are shown in the table below.

Mass of sodium carbonate in g	**Volume of carbon dioxide gas in cm^3**
0.07	16.0
0.12	27.5
0.23	52.0
0.29	12.5
0.34	77.0
0.54	95.0
0.59	95.0
0.65	95.0

The result for 0.29 g of sodium carbonate is anomalous.

Suggest what may have happened to cause this anomalous result.

_____

_____

(1)

(c) Why does the volume of carbon dioxide collected stop increasing at 95.0 cm^3?

_____

_____

(1)

(d) What further work could the student do to be more certain about the minimum mass of sodium carbonate needed to produce 95.0 cm^3 of carbon dioxide?

_____

_____

(1)

(e) The carbon dioxide was collected at room temperature and pressure.
The volume of one mole of any gas at room temperature and pressure is 24.0 dm^3.

How many moles of carbon dioxide is 95.0 cm^3?

Give your answer in three significant figures.

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_____

_____

_____

_____

_____ mol

(2)

(f) Suggest **one** improvement that could be made to the apparatus used that would give more accurate results.

Give a reason for your answer.

_____

_____

_____

_____

(2)

(g) One student said that the results of the experiment were wrong because the first few bubbles of gas collected were air.

A second student said this would make no difference to the results.

Explain why the second student was correct.

_____

_____

_____

_____

_____

_____

(2)

(Total 11 marks)

Q11.

Metals are extracted from ores in the Earth's crust.

Some ores contain metal carbonates and some ores contain metal oxides.

(a) (i) Name the type of reaction that happens when a metal carbonate is heated.

_____

(1)

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(ii) Which solid product is formed when copper carbonate is heated?

Tick (✔) **one** box.

copper ☐

copper nitrate ☐

copper oxide ☐

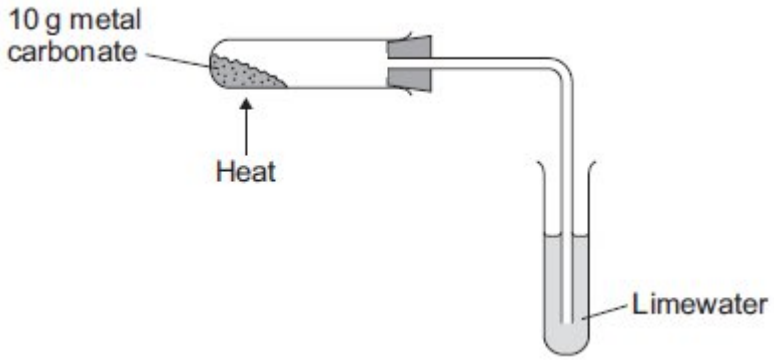
copper sulfide ☐

(1)

(b) A student investigated heating four metal carbonates.

Figure 1 shows the apparatus used.

Figure 1



The student heated each metal carbonate for five minutes.

The table below shows the results.

Metal carbonate	Mass of metal carbonate at start in g	Mass of solid after heating for 5 minutes in g	Observations
Copper carbonate	10.0	6.9	Limewater turns cloudy
Magnesium carbonate	10.0	9.1	Limewater turns cloudy
Potassium carbonate	10.0	10.0	Limewater does not turn cloudy
Zinc carbonate	10.0	8.3	Limewater turns cloudy

(i) Explain the results for potassium carbonate.

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_____

_____

_____

_____

_____

_____

_____

(3)

(ii) Suggest how the reactivity series can be used to predict which metal carbonate reacts most easily when heated.

_____

_____

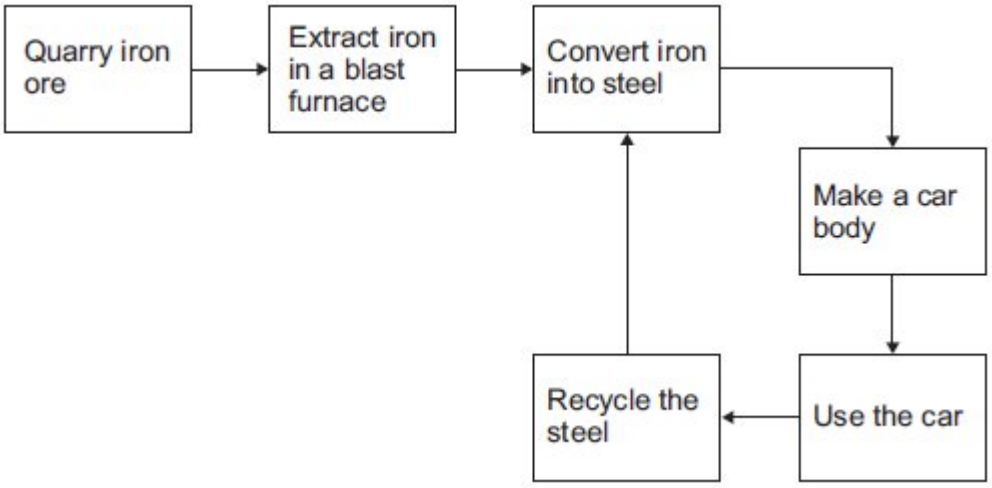
_____

_____

(2)

(c) **Figure 2** shows a simple life cycle of a car body.

Figure 2



(i) Complete the sentence.

Iron ores must contain enough iron to _____

_____

(1)

(ii) Some iron ores contain iron oxide (Fe_2O_3).

Complete and balance the equation for a reaction to produce iron from iron oxide.

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_____ Fe_2O_3 + _____ C $\longrightarrow$ _____ + _____ CO_2

(2)

(iii) Give **two** reasons why iron produced in a blast furnace is converted into steel.

_____

_____

_____

_____

(2)

(iv) When a car reaches the end of its useful life, the car body can be:

- recycled
- reused
- sent to landfill.

Give **three** reasons why a steel car body should be recycled and **not** reused or sent to landfill.

_____

_____

_____

_____

_____

_____

(3)

(Total 15 marks)

Q12.

This question is about chemical analysis.

(a) A student has solutions of three compounds, **X, Y** and **Z**.

The student uses tests to identify the ions in the three compounds.

The student records the results of the tests in the table.

	Test			
Compound	**Flame test**	**Add sodium hydroxide solution**	**Add hydrochloric acid and barium chloride**	**Add nitric acid and silver nitrate solution**

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			solution	
X	no colour	green precipitate	white precipitate	no reaction
Y	yellow flame	no reaction	no reaction	yellow precipitate
Z	no colour	brown precipitate	no reaction	cream precipitate

Identify the **two** ions present in each compound, **X, Y** and **Z**.

X _____

Y _____

Z _____

(3)

(b) A chemist needs to find the concentration of a solution of barium hydroxide. Barium hydroxide solution is an alkali.

The chemist could find the concentration of the barium hydroxide solution using two different methods.

Method 1

- An excess of sodium sulfate solution is added to 25 cm^3 of the barium hydroxide solution. A precipitate of barium sulfate is formed.
- The precipitate of barium sulfate is filtered, dried and weighed.
- The concentration of the barium hydroxide solution is calculated from the mass of barium sulfate produced.

Method 2

- 25 cm^3 of the barium hydroxide solution is titrated with hydrochloric acid of known concentration.
- The concentration of the barium hydroxide solution is calculated from the result of the titration.

Compare the advantages and disadvantages of the two methods.

_____

_____

_____

_____

_____

_____

_____

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_____

_____

_____

(5)

(Total 8 marks)

Q13.

This question is about atoms and isotopes.

(a) Atoms contain protons, neutrons and electrons.

A lithium atom has the symbol $^{7}_{3}Li$

Explain, in terms of sub-atomic particles, why the mass number of this lithium atom is 7.

_____

_____

_____

_____

_____

_____

(3)

(b) Amounts of substances can be described in different ways.

Complete the sentences.

One mole of a substance is the relative formula mass in

_____

The relative atomic mass of an element compares the mass of an atom of an element with the mass of an atom of

_____

(2)

(c) Two isotopes of oxygen are $^{18}_{8}O$ and $^{16}_{8}O$

Describe the similarities and differences between the isotopes $^{18}_{8}O$ and $^{16}_{8}O$

You should refer to the numbers of sub-atomic particles in each isotope.

_____

_____

_____

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_____

_____

_____

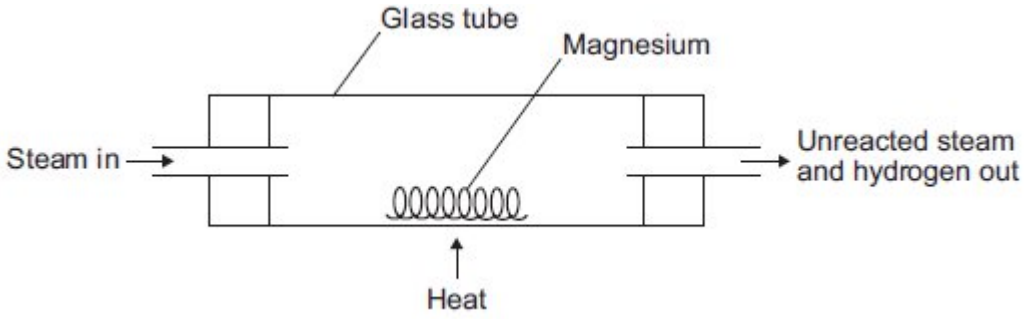
(3)

(Total 8 marks)

Q14.

Magnesium reacts with steam to produce hydrogen gas and magnesium oxide.

A teacher demonstrated the reaction to a class. The figure below shows the apparatus the teacher used.



(a) (i) The hydrogen produced was collected.

Describe how to test the gas to show that it is hydrogen.

Test _____

_____

Result _____

_____

(2)

(ii) Explain why the magnesium has to be heated to start the reaction.

_____

_____

_____

_____

(2)

(b) The equation for the reaction is:

$$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$$

(i) The teacher used 1.00 g of magnesium.

Use the equation to calculate the maximum mass of magnesium oxide produced.

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Give your answer to three significant figures.

Relative atomic masses (A_r): O = 16; Mg = 24

_____

_____

_____

_____

Maximum mass = _____ g

(3)

(ii) The teacher's demonstration produced 1.50 g of magnesium oxide.

Use your answer from part (b)(i) to calculate the percentage yield.

If you could not answer part (b)(i), use 1.82 g as the maximum mass of magnesium oxide. This is **not** the answer to part (b)(i).

_____

Percentage yield = _____ %

(2)

(iii) Give **one** reason why the percentage yield is less than 100%.

_____

_____

_____

(1)

(Total 10 marks)

Q15.

Thermosoftening polymers can be used to make plastic bottles and food packaging.

(a) Why are thermosoftening polymers **not** suitable for storing very hot food?

_____

_____

(1)

(b) The reaction to produce the polymers uses a catalyst.

Why are catalysts used in chemical reactions?

_____

_____

(1)

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(c) Compounds from food packaging must not get into food.

Gas chromatography can be used to separate compounds in food.

The output from the gas chromatography column can be linked to an instrument which can identify the compounds.

(i) Name the instrument used to identify the compounds.

_____

_____

(1)

(ii) Give **one** reason why instrumental methods of analysis are used to identify the compounds.

_____

_____

(1)

(d) Poly(ethene) is a thermosoftening polymer.

Poly(ethene) can be made with different properties. The properties depend on the conditions used when poly(ethene) is made.

Suggest **two** conditions which could be changed when poly(ethene) is made.

_____

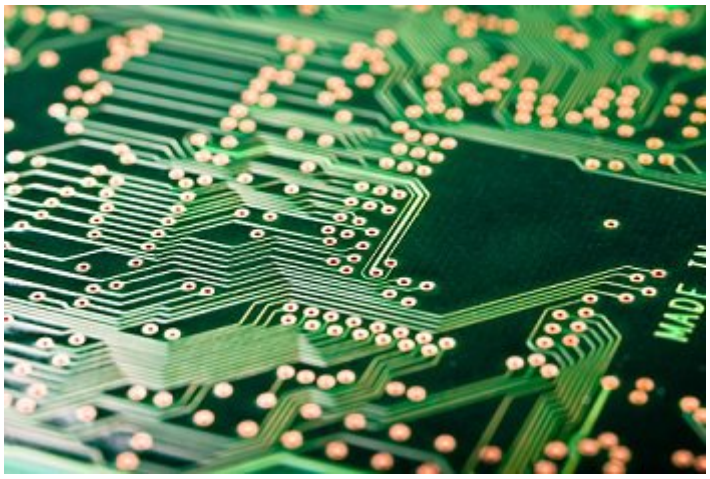
_____

(2)

(Total 6 marks)

Q16.

Etching is a way of making printed circuit boards for computers.



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Printed circuit boards are made when copper sheets are etched using iron(III) chloride solution. Where the copper has been etched, only plastic remains.

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(a) Copper is a good conductor of electricity.

Explain why.

_____

_____

_____

_____

(2)

(b) Iron(III) chloride can be produced by the reaction shown in the equation:

$$2\ Fe + 3\ Cl_2 \rightarrow 2\ FeCl_3$$

(i) Calculate the maximum mass of iron(III) chloride ($FeCl_3$) that can be produced from 11.20 g of iron.

Relative atomic masses (A_r): Cl = 35.5; Fe = 56.

_____

_____

_____

_____

_____

Maximum mass of iron(III) chloride = _____ g

(3)

(ii) The actual mass of iron(III) chloride ($FeCl_3$) produced was 24.3 g.

Calculate the percentage yield.

(If you did not answer part (b)(i) assume that the maximum theoretical mass of iron(III) chloride ($FeCl_3$) is 28.0 g. This is **not** the correct answer to part (b)(i).)

_____

_____

Percentage yield = _____%

(1)

(Total 6 marks)

Q17.

This question is about lithium and sodium.

(a) Use the Chemistry Data Sheet to help you to answer this question.

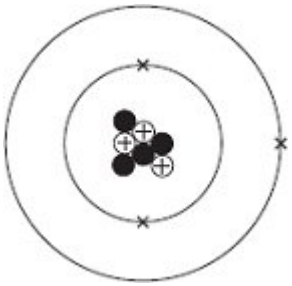
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In which group of the periodic table are lithium and sodium? Group ☐

(1)

(b) A lithium atom can be represented as $^{7}_{3}Li$

The diagram represents the lithium atom.



(i) Some particles in the nucleus have a positive charge.

What is the name of these particles?

_____

(1)

(ii) Some particles in the nucleus have no charge.

What is the name of these particles?

_____

(1)

(iii) Use the correct answer from the box to complete the sentence.

3	**4**	**7**

The mass number of this atom of lithium is ☐

(1)

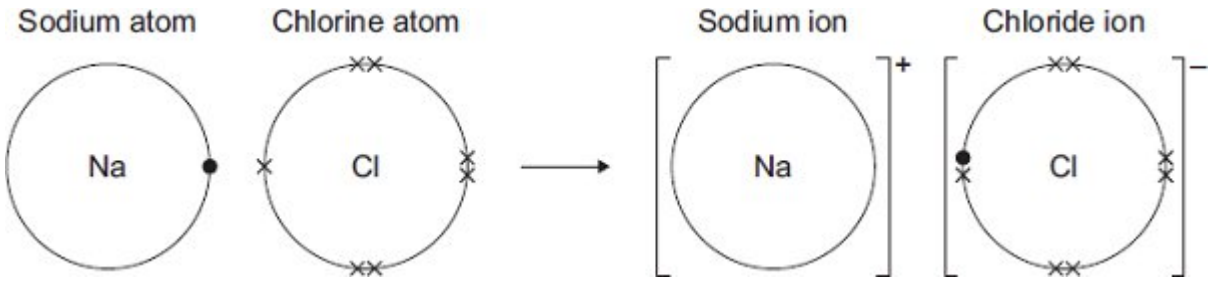
(c) Sodium reacts with chlorine to produce sodium chloride.

sodium + chlorine ⟶ sodium chloride

The diagram shows how the reaction happens.

Only the outer electrons are shown.

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Draw a ring around the correct answer to complete each sentence.

(i) A sodium atom changes into a sodium ion by [gaining / losing / sharing] an electron.

(1)

(ii) A sodium ion has [a negative / no / a positive] charge.

(1)

(iii) The ions in sodium chloride are held together by strong [covalent / electrostatic / magnetic] forces.

(1)

(d) Sodium chloride is an ionic compound.

Tick (✓) **two** properties of ionic compounds.

Property	Tick (✓)
Do **not** dissolve in water	
High melting points	
Low boiling points	
Strong bonds	

(2)

(e) (i) The formula of sodium chloride is NaCl

Calculate the relative formula mass of sodium chloride.

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Relative atomic masses: Na = 23; Cl = 35.5

_____

_____

Relative formula mass = _____

(1)

(ii) Draw a ring around the correct answer to complete each sentence.

The relative formula mass of a substance, in grams,

is one | ion / isotope / mole | of the substance.

(1)

(f) Nanoparticles of sodium chloride (salt) are used to flavour crisps.

What are nanoparticles?

_____

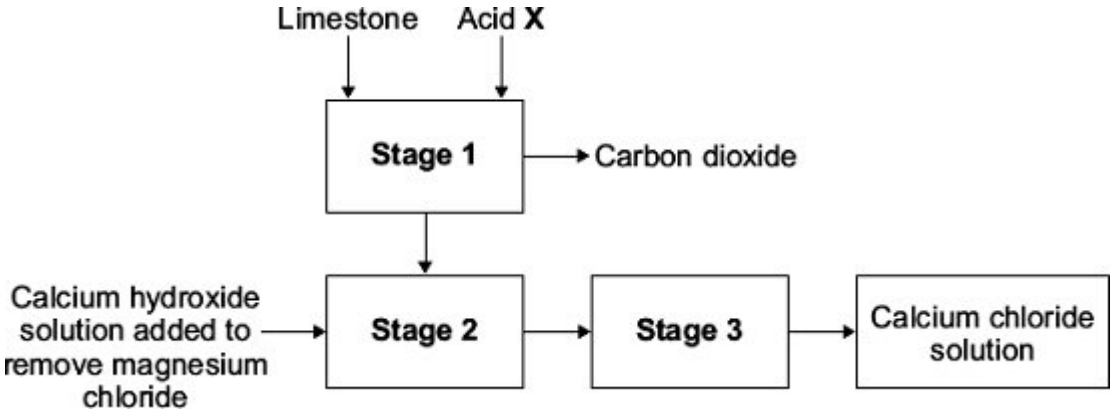
_____

(1)

(Total 12 marks)

Q18.

(a) Calcium chloride is made from limestone. Limestone contains mainly calcium carbonate and a small amount of magnesium carbonate.



(i) In **stage 1** calcium carbonate reacts with acid **X** to form calcium chloride.

Draw a ring around the name of acid **X**.

hydrochloric **nitric** **sulfuric**

(1)

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(ii) **Stage 1** produces a concentrated solution of calcium chloride.
The solution also contains magnesium chloride.

Calcium hydroxide solution is added in **stage 2** to remove the magnesium chloride.

The equation for this reaction is:

$$MgCl_2(aq) + Ca(OH)_2(aq) \rightarrow Mg(OH)_2(s) + CaCl_2(aq)$$

Draw a ring around the correct answer to complete each sentence.

In **stage 2** a precipitate is made because

magnesium hydroxide is [dissolved / insoluble / soluble] in water.

In **stage 3** the solid magnesium hydroxide can be separated from the calcium chloride solution using [chromatography. / electrolysis. / filtration.]

(2)

(iii) What method can be used to change the calcium chloride solution into solid calcium chloride?
Draw a ring around your answer.

crystallisation **electrolysis** **reduction**

(1)

(b) Calcium chloride can also be made by reacting calcium with chlorine:

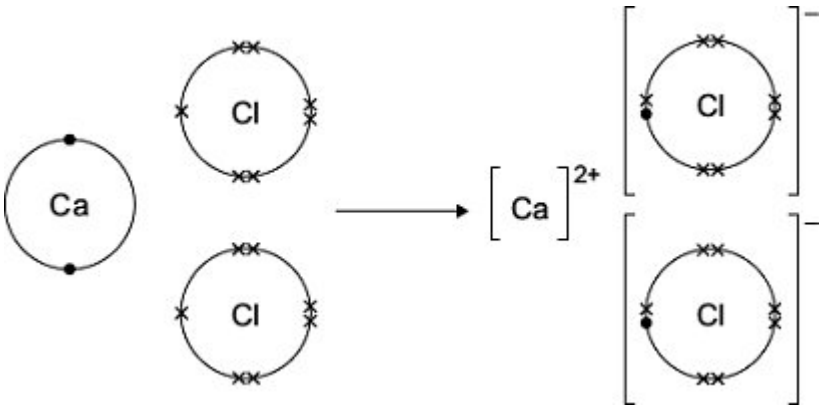
calcium + chlorine → calcium chloride

The diagram shows what happens to atoms of calcium and chlorine in this reaction.

The dots (●) and crosses (x) are used to represent electrons.

Only the outer electrons are shown.

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Use the diagram to help you to answer this question.

Describe, as fully as you can, what happens when calcium reacts with chlorine to make calcium chloride.

_____

_____

_____

_____

_____

_____

_____

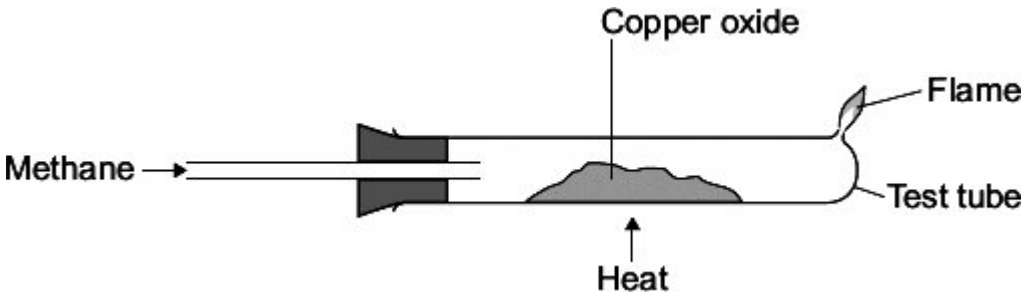
_____

(4)

(Total 8 marks)

Q19.

An experiment was done on the reaction of copper oxide (CuO) with methane (CH_4).



(a) The equation for this reaction is shown below.

$$4CuO(s) + CH_4(g) \rightarrow 4Cu(s) + 2H_2O(g) + CO_2(g)$$

The water and carbon dioxide produced escapes from the test tube.

Use information from the equation to explain why.

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_____

(1)

(b) (i) Calculate the relative formula mass (M_r) of copper oxide (CuO).

Relative atomic masses (A_r): O = 16; Cu = 64.

_____

_____

_____

Relative formula mass (M_r) = _____

(2)

(ii) Calculate the percentage of copper in copper oxide.

_____

_____

_____

Percentage of copper = _____ %

(2)

(iii) Calculate the mass of copper that could be made from 4.0 g of copper oxide.

_____

_____

Mass of copper = _____ g

(1)

(c) The experiment was done three times.
The mass of copper oxide used and the mass of copper made was measured each time.
The results are shown in the table.

	Experiment		
	1	**2**	**3**
Mass of copper oxide used in g	4.0	4.0	4.0
Mass of copper made in g	3.3	3.5	3.2

(i) Calculate the mean mass of copper made in these experiments.

_____

_____

Mean mass of copper made = _____ g

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(1)

(ii) Suggest how the results of these experiments could be made more precise.

_____

_____

(1)

(iii) The three experiments gave slightly different results for the mass of copper made.
This was caused by experimental error.

Suggest **two** causes of experimental error in these experiments.

1. _____

_____

2. _____

_____

(2)

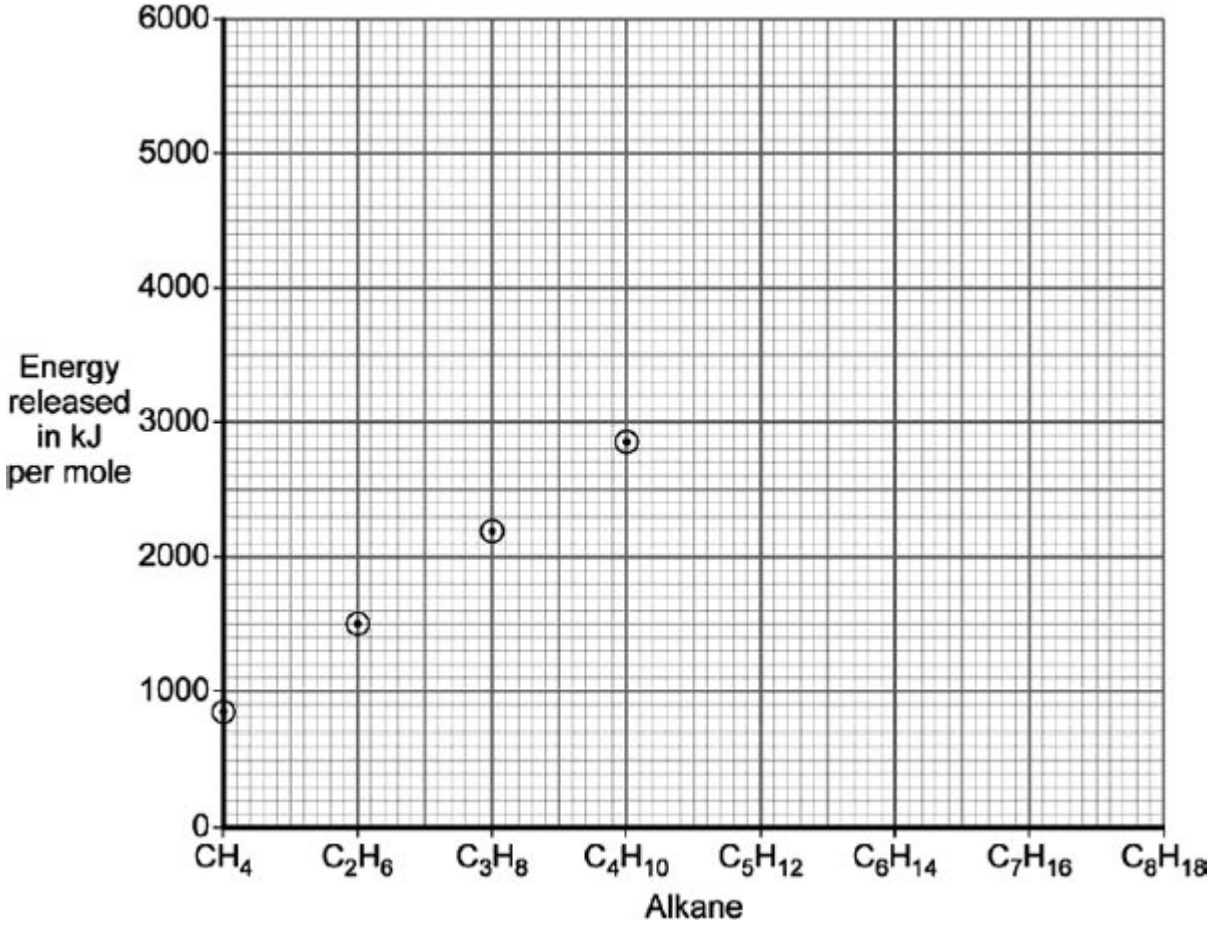
(Total 10 marks)

Q20.

(a) Alkanes are important hydrocarbon fuels. They have the general formula C_nH_{2n+2}

The points on the graph show the amount of energy released when 1 mole of methane (CH_4), ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}) are burned separately.

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(i) Draw a line through the points and extend your line to the right-hand edge of the graph.

(1)

(ii) Use the graph to estimate the amount of energy released when 1 mole of octane (C_8H_{18}) is burned.

Energy released = _____ kJ

(1)

(iii) Suggest why we can make a good estimate for the energy released by 1 mole of pentane (C_5H_{12}).

_____

_____

(1)

(iv) A student noticed that octane (C_8H_{18}) has twice as many carbon atoms as butane (C_4H_{10}), and made the following prediction:

"When burned, 1 mole of octane releases twice as much energy as 1 mole of butane."

Use the graph to decide if the student's prediction is correct. You **must** show your working to gain credit.

_____

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_____

_____

_____

(2)

(b) Some information about four fuels is given in the table.

			Combustion products			
Fuel	**Type**	**Heat released in kJ per g**	**CO_2**	**SO_2**	**H_2O**	**Type of flame**
Bio-ethanol	Renewable	29	✓		✓	Not smoky
Coal	Non-renewable	31	✓	✓	✓	Smoky
Hydrogen	Renewable	142			✓	Not smoky
Natural gas	Non-renewable	56	✓		✓	Not smoky

From this information a student made two conclusions.

For each conclusion, state if it is correct **and** explain your answer.

(i) "Renewable fuels release more heat per gram than non-renewable fuels."

_____

_____

_____

_____

(2)

(ii) "Non-renewable fuels are better for the environment than renewable fuels."

_____

_____

_____

_____

_____

_____

_____

(2)

(Total 9 marks)

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